Unit 1: Quantities and Units

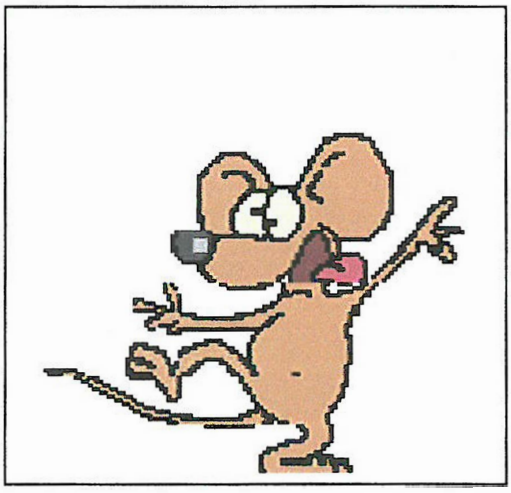
Measurement:

Much of the chemistry we do will be quantitative, requiring measured, numbered quantities. All measurements require a number and a unit associated with it to be meaningful. All the units we use will be metric, based on the SI system. The SI system, in place since 1960, has seven fundamental base units from which all other units are derived. They are

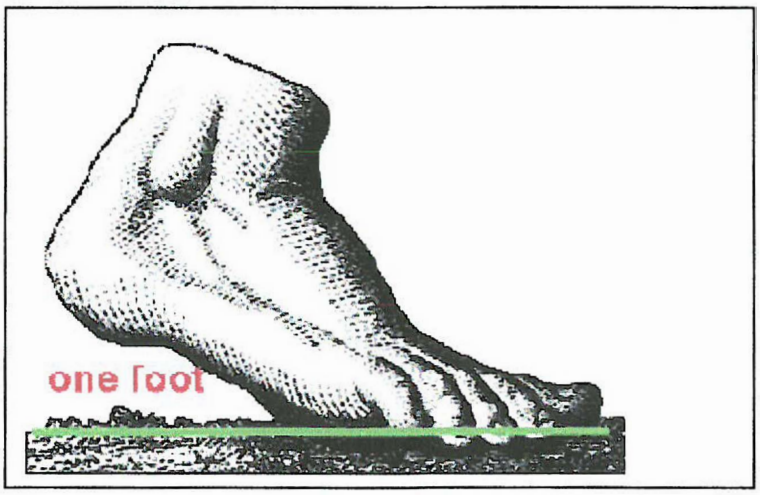
Name	Quantity	Symbol
Mass	**(1 kg = 2.2 pounds)**	**(kg) kilogram or (g) gram**
Length	**(1 m = 1.0936 yd)**	**(m) meter**
Time	**(3600s = 1Hour)**	**(s) second**
Temperature		**(K) Kelvin**
Current		**(A) Ampere**
Luminous intensity		**(cd) candela**
Amount of a substance	**(1mol=6.023x10^{23}molcules)**	**(mol) mole**

Table 5. SI prefixes

Factor	Name	Symbol	Factor	Name	Symbol
10^{24}	yotta	Y	10^{-1}	deci	d
10^{21}	zetta	Z	10^{-2}	centi	c
10^{18}	exa	E	10^{-3}	milli	m
10^{15}	peta	P	10^{-6}	micro	μ
10^{12}	tera	T	10^{-9}	nano	n
10^{9}	giga	G	10^{-12}	pico	p
10^{6}	mega	M	10^{-15}	femto	f
10^{3}	kilo	k	10^{-18}	atto	a
10^{2}	hecto	h	10^{-21}	zepto	z
10^{1}	deka	da	10^{-24}	yocto	y



The English system of measurement grew out of the creative way that people measured for themselves. Familiar objects and parts of the body were used as measuring devices. For example, people measured shorter distances on the ground with their feet.



Conversion Tables

Metric System of Measurements

Length

10 millimeters = 1 centimeter
10 centimeters = 1 decimeter
10 decimeters = 1 meter
10 meters = 1 decameter
10 decameters = 1 hectometer
10 hectometers = 1 kilometer
1000 meters = 1 kilometer
1 mile = 5280ft = 1.609×10^5cm = 1.609×10^3m = 1.609km
1foot = 12in 1.0in = 2.54cm = 0.0254m = 25.4mm

Area

100 mm^2	= 1 cm^2
10 000 cm^2	= 1 m^2
100 m^2	= 1 are
100 ares	= 1 hectare
10 000 m^2	= 1 hectare
100 hectares	= 1 km^2
1000000 m^2	= 1 km^2
1yd	= 36in

Volume

1000 mm^3	= 1 cm^3
1000 cm^3	= 1 dm^3
1000 dm^3	= 1 m^3
1 million cm^3	= 1 m^3

1 Liter = 1dm^3 = $10^{-3}m^3$

Capacity

10 mL = 1 cL
10 cL = 1 dL
10 dL = 1 L
1000 L = 1 m^3

Mass

1000 g = 1 kg
1000 kg = 1 ton
10000000g = 1 metric ton
1 pound = 0.4536kg
1US Metric Ton = 907Kg

The U K (Imperial) System of Measurements

Length

12 inches = 1 foot
3 feet = 1 yard
22 yards = 1 chain
10 chains = 1 furlong
8 furlongs = 1 mile
5280 feet = 1 mile
1760 yards = 1 mile

Area

144 sq. inches = 1 square foot
9 sq. feet = 1 square yard
4840 sq. yards = 1 acre
640 acres = 1 square mile

Volume

1728 in^3 = 1 ft^3
27 ft^3 = 1 yd^3
1L = 1 dm^3
1 L = 0.01 m^3

Capacity

20 fluid ounces = 1 pint
4 gills = 1 pint
2 pints = 1 quart
4 quarts = 1 gallon (8 pints)

Mass (Avoirdupois)

437.5 grains = 1 ounce
16 ounces = 1 pound (7000 grains)
14 pounds = 1 stone
8 stones = 1 hundredweight [cwt]
20 cwt = 1 ton (2240 pounds)

Troy Weights

24 grains = 1 pennyweight
20 pennyweights = 1 ounce (480 grains)
12 ounces = 1 pound (5760 grains)

Apothecaries' Measures

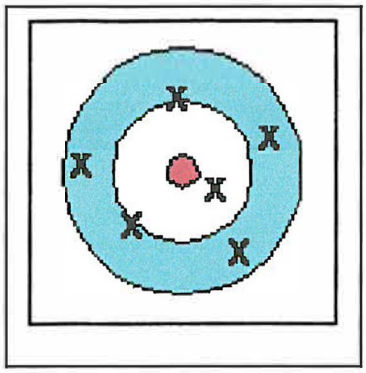
20 minims = 1 fl.scruple
3 fl.scruples = 1 fl.drachm
8 fl.drachms = 1 fl.ounce
20 fl.ounces = 1 pint

Apothecaries' Weights

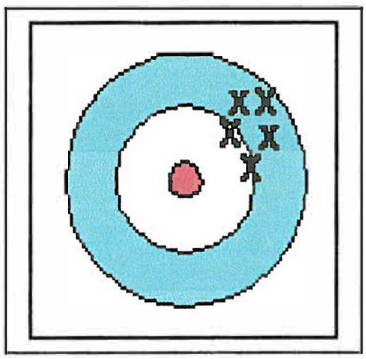
20 grains = 1 scruple
3 scruples = 1 drachm
8 drachms = 1 ounce (480 grains)
12 ounces = 1 pound (5760 grains)

PRECISION VERSUS ACCURACY

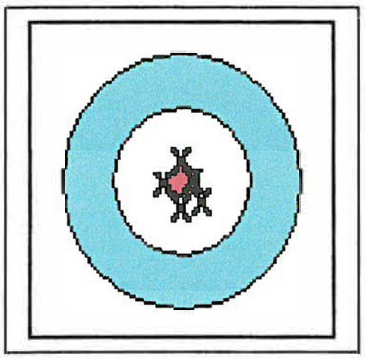
Accuracy refers to how closely a measured value agrees with the correct value.
Precision refers to how closely individual measurements agree with each other.



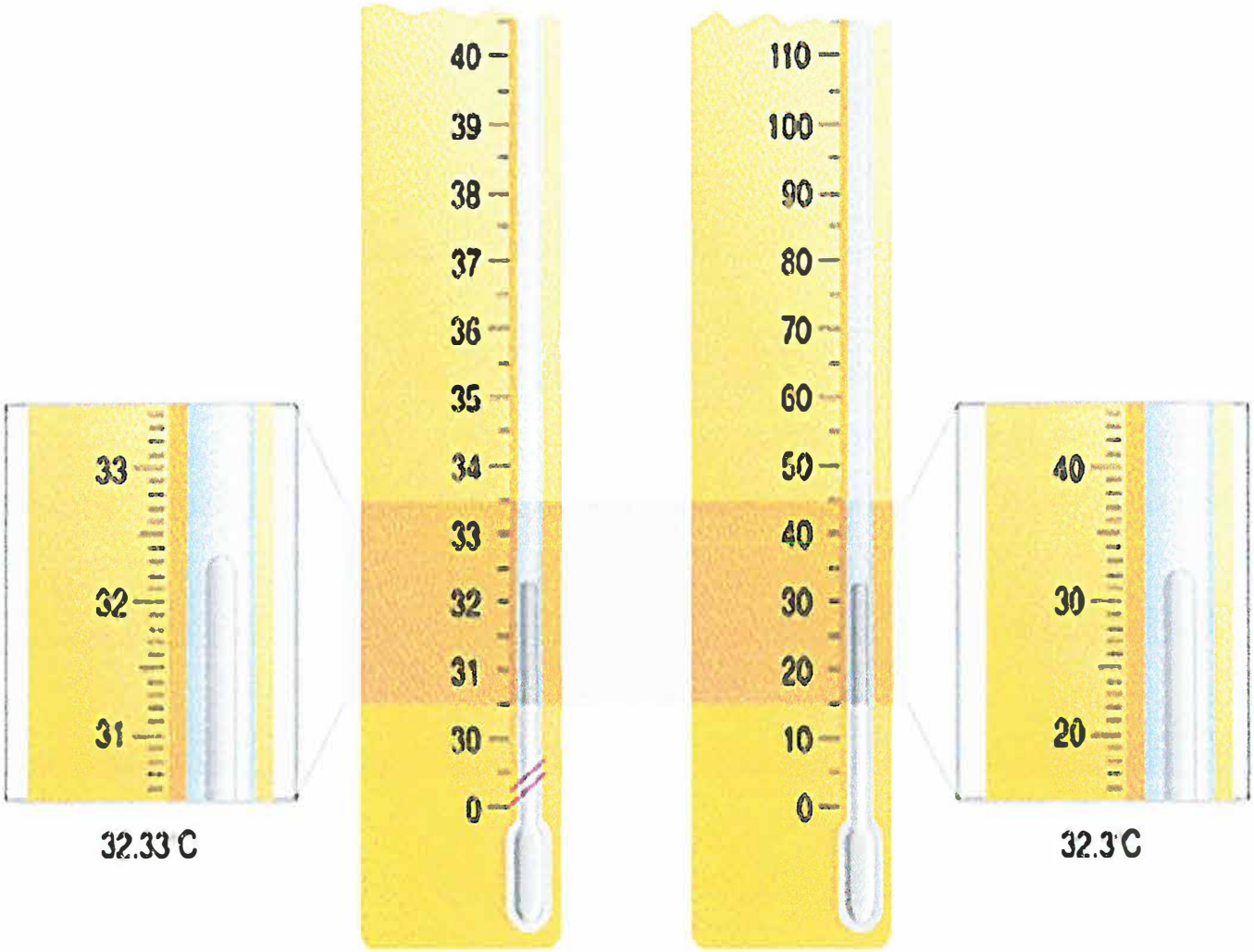
Accurate: (the average is accurate) not precise



Precise, not accurate



Accurate and Precise



Scientific Notation

Chemistry deals with very large and very small numbers. Consider this calculation:

(0.000000000000000000000000000000663 x 30,000,000,000) ÷ 0.00000009116

Hopefully you can see, how awkward it is. Try keeping track of all those zeros! In scientific notation, this problem is:

$$(6.63 \times 10^{-31} \times 3.0 \times 10^{10}) \div 9.116 \times 10^{-8}$$

It is now much more compact, it better represents significant figures, and it is easier to manipulate mathematically. The trade-off, of course, is that you have to be able to read scientific notation.

Format for Scientific Notation

1. Used to represent positive numbers only.

2. Every positive number X can be written as: $(1 < N < 10) \times 10^{\text{some positive or negative integer}}$

Where N represents the numerals of X with the decimal point after the <u>first</u> nonzero digit.

3. A decimal point is in standard position if it is behind the first non-zero digit. Let X be any number and let N be that number with the decimal point moved to standard position.

- **Rules to consider:**
- **<u>The first number must be a number that is ≥ 1 but < 10</u>**
- **If the decimal point on any whole number or decimal is moved to the:**
  - **Right, the exponent will be a negative sign**
  - **Left, the exponent will be a positive sign**

4. Some examples of number three:
- 0.00087 becomes 8.7×10^{-4}
- 9.8 becomes 9.8×10^{0} (the 10^{0} is seldom written)
- 23,000,000 becomes 2.3×10^{7}

5. Some more examples of number three:
- 0.000000809 becomes 8.09×10^{-7}
- 4.56 becomes 4.56×10^{0}
- 250,000,000,000 becomes 2.50×10^{11}

Example #1 - Convert 29,190,000,000 to scientific notation.

First Explanation

Step 1 - start at the decimal point of the original number and count the number of decimal places you move, stopping to the right of the first non-zero digit. Remember that's the first non-zero digit counting from the left.

Step 2 - The number of places you moved (10 in this example) will be the exponent. If you moved to the left, it's a positive value. If you moved to the right, it's negative.

The answer is 2.919×10^{10}.

Example 2 - Write 0.00000000459 in scientific notation.

Step 1 - Write all the significant digits down with the decimal point just to the right of the first significant digit. Like this: 4.59. Please be aware that this process should ALWAYS result in a value between 1 and 10.

Step 2 - Now count how many decimal places you would move from 4.59 to recover the original number of 0.00000000459. The answer in this case would be 9 places to the LEFT. That is the number 0.000000001. Be aware that this number in exponential notation is 10^{-9}.

Step 3 - Write 4.59 times the other number, BUT, write the other number as a power of 10. The number of decimal places you counted gives the power of ten. In this example, that power would be 9. The correct answer to this step is: 4.59×10^{-9}

Suppose the number to be converted looks something like scientific notation, but it really is not. For example, look carefully at the example below. Notice that the number 428.5 is not a number between 1 and 10. Although writing a number in this fashion is perfectly OK, it is not in standard scientific notation. What would it look like when converted to standard scientific notation?

Example #3 - Convert 428.5×10^{9} to scientific notation.

Step 1 - convert the 428.5 to scientific notation. (The lesson up to this point has been covering how to do just this step). Answer = 4.285×10^{2}.

Step 2 - write out the new number. Answer = $4.285 \times 10^{2} \times 10^{9}$.

Step 3 - combine the exponents according to the usual rules for exponents. Answer = 4.285×10^{11}.

Example #4 - convert 208.8×10^{-11} to scientific notation.

Step 1 - convert the 208.8 to scientific notation. Answer = 2.088×10^{2}.

Step 2 - write out the new number. Answer = $2.088 \times 10^{2} \times 10^{-11}$.

Step 3 - combine the exponents according to the usual rules for exponents. Answer = 2.088×10^{-9}.

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1. When converting a number greater than one (the 428.5 and the 208.8 in the previous examples), the resulting exponent will become more positive (11 is more positive than 9 while -9 is more positive than -11).

2. When converting a number less than one (the 0.000531 and the 0.00000306 in the previous examples), the resulting exponent will always be more negative (10 is more negative than 14 and -23 is more negative than -17).

If the decimal point is moved to the left, the exponent goes up in value (becomes more positive).
If the decimal point is moved to the right, the exponent goes down in value (becomes more negative).

Significant Figures: The easy way!

There are three rules on determining how many significant figures are in a number:

1. Non-zero digits are always significant.
2. Any zeros between two significant digits are significant.
3. A final zero or trailing zeros in the decimal portion ONLY are significant.

Whole Numbers: Determining the number of sig-figs without a decimal showing.

- 125052 **- 6 sig-figs, the zero in between two numbers count.**
- 6500 **- 2 sig-figs, the trailing zeros do not count**
- 18540000 **- 4 sig-figs, the trailing zeros do not count**

Whole numbers with decimals:

- 5.417 **- 4 sig-figs, all numbers greater than zero count.**
- 40.06 **- 4 sig-figs, zeros between non-zero digits are significant.**
- 1600.0 **- 5 sig-figs, final zeros to the right of the decimal point are significant**

Decimals:

- 0.05 **- 1 sig-fig, the zeros mark the position of the decimal.**
- 0.060 **- 2 sig-figs, the first two mark position, the final zero is significant.**
- 0.00682 **- 3 sig-fig, the zeros mark the position of the decimal.**

Scientific Notation:

- 3×10^3 **- 1 sig-fig**
- 5.4×10^5 **- 2 sig-figs**
- 3.000×10^3 **- 4 sig-figs, final zeros to the right of the decimal point are significant**

Temperature Conversion and Density

Formulas:

Temperature

$$K = C^0 + 273.15$$

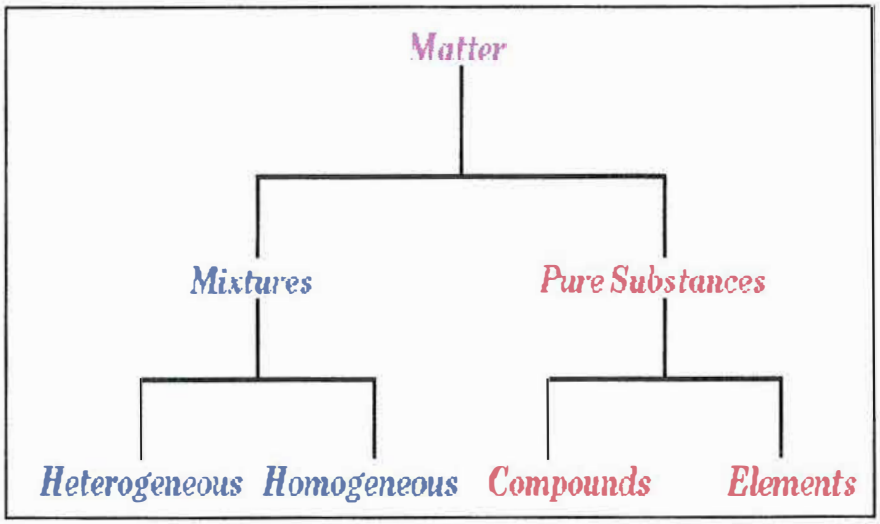
$$C^0 = \frac{(F^0 - 32)}{1.8}$$

$$F^0 = (1.8)(C^0) + 32$$

Density

$$D = \frac{\text{Mass}}{\text{Volume}}$$

Classification of Matter



- **Physical change**

A change in the form of a substance, for instance, from solid to liquid or liquid to gas or solid to gas, without changing the chemical composition of the substance. As we will see later, chemical bonds are not broken in a physical change.
Examples: **Boiling of water and the melting of ice.**

- **Chemical change**

The change of a substance into another substance, by reorganization of the atoms, i.e. by the making and breaking of chemical bonds. In a chemical change, a chemical reaction takes place.
Examples: **Rusting of iron and the decomposition of water, induced by an electric current, to gaseous hydrogen and gaseous oxygen.**

- **Mixture**

Two or more substances, combined in varying proportions - each retaining its own specific properties. The components of a mixture can be separated by physical means, i.e. without the making and breaking of chemical bonds.
Examples: **Air, table salt thoroughly dissolved in water, milk, wood, and concrete.**

- **Heterogeneous Mixture**

Mixture in which the properties and composition are not uniform throughout the sample.
Examples: **Milk, wood, and concrete.**

- **Homogeneous Mixture**

Mixture in which the properties and composition are uniform throughout the sample. Such mixtures are termed solutions.
Examples: **Air and table salt thoroughly dissolved in water.**

- **Pure Substance**

A substance with constant composition. Can be classified an either an element or as a compound.
Examples: **Table salt (sodium chloride, NaCl), sugar (sucrose, $C_{12}H_{22}O_{11}$), water (H_2O), iron (Fe), copper (Cu), and oxygen (O_2).**

- **Element**

A substance that cannot be separated into two or more substances by ordinary chemical (or physical) means. We use the term ordinary chemical means to exclude nuclear reactions. Elements are composed of only one kind of atom. ***Examples:*** **Iron (Fe), copper (Cu), and oxygen (O_2).**

- **Compound**

A substance that contains two or more elements, in definite proportion by weight. The composition of a pure compound will be invariant, regardless of the method of preparation. Compounds are composed of more than one kind of atom. The term molecule is often used for the smallest unit of a compound that still retains all of the properties of the compound. *Examples:* Table salt (sodium chloride, NaCl), sugar (sucrose, $C_{12}H_{22}O_{11}$), and water (H_2O).

- **Conservation of Mass**

Usually attributed to Lavoisier (in 1789). "Matter (mass) is neither created nor destroyed". In other words, in a closed system (nothing escapes), any process will not change the total "matter content" (i.e. mass) of the system.

- **Law of Definite Proportion or Constant Composition**

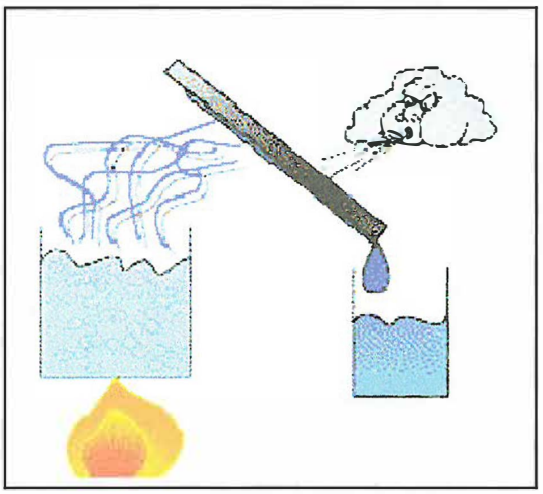
Usually attributed to Dalton and/or Proust (circa 1808). "Regardless of the method of separation, a pure compound will always contain the same elements, in the same proportion by mass." Dalton's specific contribution to this law is with regard to the inclusion of the consequences of the atomic hypothesis (Democritus, circa 400 B.C.). Dalton reasoned that since these elements were indivisible atoms, "each pure compound should contain the same proportion of these atoms, regardless of the method of preparation."

- **The proportion by mass of a species in a sample can be expressed as a mass percent (%). The definition is:**

$$\text{Mass \%} = \frac{\text{Part of mass X}}{\text{Total mass of sample}} \times 100\%$$

Separating Chemical Components &Methods in Chemistry

- **Distillation: a process in which a liquid or vapor mixture of two or more substances is separated into its component fractions of desired purity, by the application and removal of heat.**



- **Distillation is the most common separation technique.**
- **It consumes enormous amounts of energy, both in terms of cooling and heating requirements.**
- **It can contribute to more than 50% of plant operating costs**

- **Filtration: a mixture is poured onto a mesh, such as filter paper, which passes the liquid and leaves the solid behind.**
- **Chromatography: The components to be separated are distributed between two phases: a *stationary phase* (solid) bed and a *mobile phase (liquid)*, which percolates through the stationary bed. The separation occurs due to different affinities for the two phases, & thus moves at different rates.**

The Periodic Table

J.A.R. Newlands - 1867 first version of Periodic Table. Newlands arranged the known elements by increasing atomic mass along horizontal rows seven elements long, stated that the 8th element would have similar properties to the first from the series. Newlands called this the law of octaves. Newlands' work failed after Ca in predicting a consistent trend.

Dimitri Mendeleev 1869, Professor of Chemistry at the University of Saint Petersburg (Leningrad). Mendeleev stated that the elements vary periodically (in cycles) according to their atomic masses.

Mendeleev separated his elements and left spaces on his table in order for the periodicity to continue. He then predicted that elements would be discovered to fill these "gaps" in the table. Mendeleev even accurately stated the properties of these elements. Scandium (ekaboron), galluim (ekaaluminum), and germanium (ekasilicon). By 1886, all of the elements predicted by Mendeleev had been isolated.

When Mendeleev's notes show that the periodic system was created in a single day, February 17, 1869. He arrived at his system by puzzling over cards containing the names of the 63 known elements along with their atomic weights and important chemical and physical properties.

Breakdown of the Periodic Table…

- Group 1 = **Alkali Metals** – are very reactive metals that do not occur freely in nature. Alkali metals can explode if they are exposed to water.

- Group 2 = **Alkali Earth Metals** – alkaline metals are not found free in nature.

- Group 3 to 12 = **Transition Metals** – are both ductile and malleable, and conduct electricity and heat. Three important elements; iron, cobalt, and nickel, and they are the only elements known to produce a magnetic field.

- Group 13 to 15 = **Other Metals** (Aluminum, Gallium, Indium, Tin, Thallium, Lead and Bismuth) – These elements do not exhibit variable oxidation states, and their valence electrons are only present in their outer shell. All of these elements are solid, have relatively high densities.

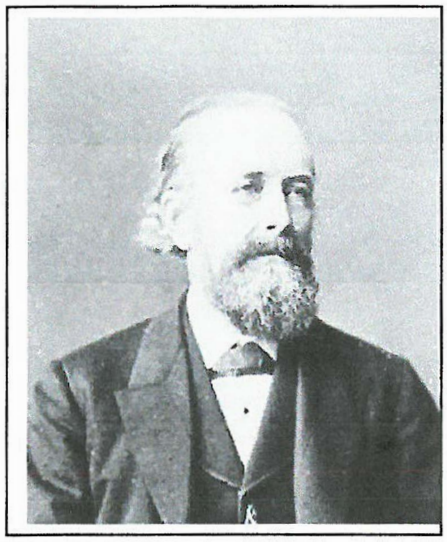
- Group 13 to 15 = **Metalloids** (Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium and Polonium) - Metalloids have properties of both metals and non-metals. Some of the metalloids, such as silicon and germanium, are semi-conductors. This property makes metalloids useful in computers and calculators.

- Group 14 to 16 = **Non-Metals** (Hydrogen, Carbon, Nitrogen, Oxygen, Phosphorus, Sulfur, Selenium – Non-Metals are not able to conduct electricity or heat very well.

- Group 17 = **Halogens** – are the five non-metallic elements. The term "halogen" means "salt-former" and compounds containing halogens are called "salts". The halogens exist, at room temperature, in all three states of matter:

  Solid- Iodine, Astatine
  Liquid- Bromine
  Gas- Fluorine, Chlorine

- Group 18 = **Noble Gases** – All noble gases have the maximum number of electrons possible in their outer shell (2 for Helium, 8 for all others), making them stable.



Why he was famous?

A pioneering chemist, Frankland was to 'invent' the chemical bond, and became known as the father of Valency. He also invented the science of organo-metallic chemistry, which is the study of compounds of metal with groups of atoms

Formulas and Names of Compounds

- **Chemical Formula -** is a combination of symbols that represent the composition of a compound.
- **Binary Compounds -** compounds that contain only two elements.
- **Polyatomic Compounds -** compounds that contain many atoms or elements and contains more than two covalently bound atoms.
- **Oxidation Number -** a positive or negative number that indicates an element's ability to form a compound.
- **Anion -** An anion is a negatively charged ion. Nonmetals typically form anions.
- **Cation -** A cation is a positively charged ion. Metals typically form cations.

Steps for writing formulas: example mixing sulfur and aluminum

1. Write the symbol of the positive element followed by the symbol of the negative element.

   Al S

2. Look up the Oxidation numbers for each element. Write the Oxidation number above each element.

   Al^{+3} S^{-2}

3. Take the Oxidation numbers and move them in a crisscross pattern to the opposite element.

Al^{+3} S^{-2} Therefore Al_2S_3 - Aluminum Sulfide

Examples:

1. Magnesium + Chlorine
   Mg^{+2} Cl^{-1}
   Mg_1Cl_2 or $MgCl_2$

2. Hydrogen + Sulphur
   H^{+1} S^{-2}
   H_2S_1 or H_2S

3. Cesium + Bromine
   Cs^{+2} Br^{-1}
   Cs_1Br_2 or $CsBr_2$

4. Potassium + Iodine
   K^{+1} I^{-1}
   K_1I_1 or KI

Naming Binary Compounds using the Chemical Name (IUPAC)

- International Union of Pure and Applied Chemistry - Latin and Greek stems are used to represent the elements with a universal meaning.

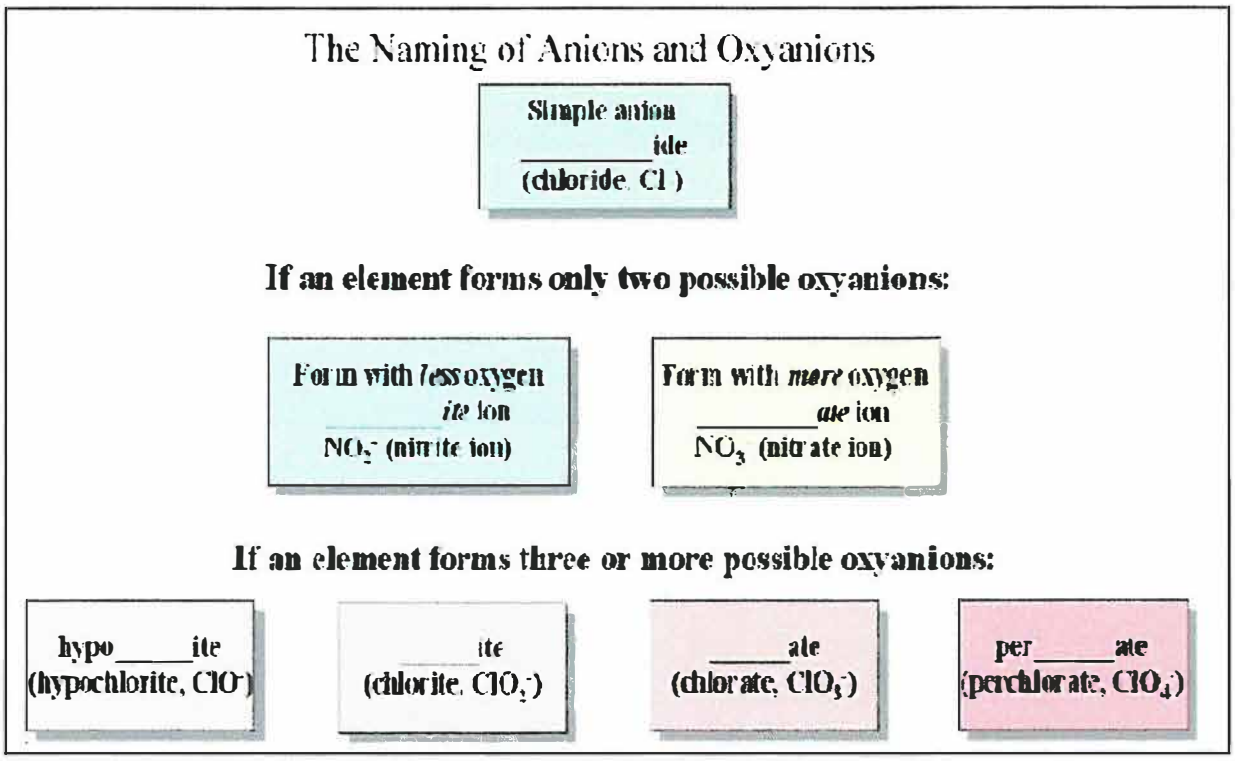
Prefix:	Number of Atoms:
Mono-	One (1)
Di -	Two (2)
Tri -	Three (3)
Tetra -	Four (4)
Penta -	Five (5)
Hexa -	Six (6)
Hepta -	Seven (7)
Octa -	Eight (8)
Nona -	Nine (9)
Deca -	Ten (10)

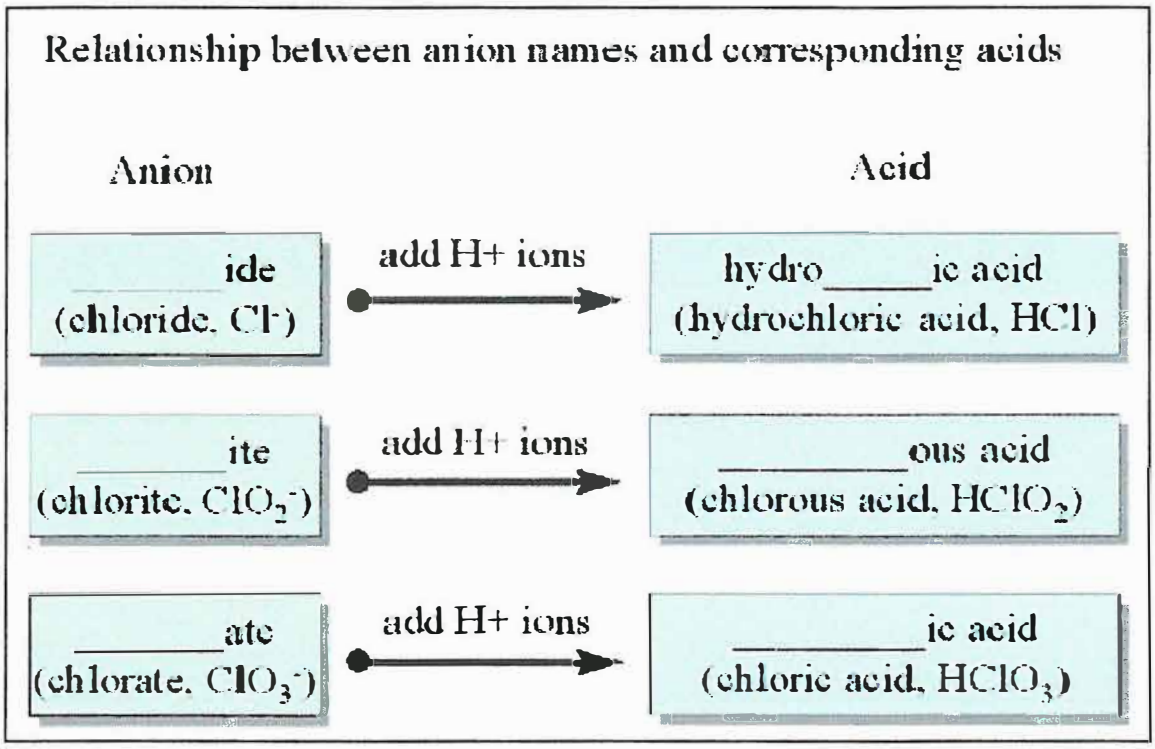
Suffix ending: When you have a binary compound, the suffix ending ends in "ide".

One Important Rule: The prefix "Mono" may be omitted for the first element. The absence of a prefix for the first element usually implies that there is only one atom of that element present in the molecule.

Molecular Compounds:

Symbol:	Chemical Name:	Common Name:
HCl	**Hydrogen Chloride**	**Hydrochloric Acid**
HBr	**Hydrogen Bromide**	**Hydrobromic Acid**
NH_3	**Nitrogen Trihydride**	**Ammonia**
H_2O	**Dihydrogen Oxide**	**Water**
NaCl	**Sodium Chloride**	**Table Salt**





Exception: With acids involving sulfur, like hydrogen sulfate, add a <u>ur-</u> before the suffix.

<u>Examples:</u>

a) HNO_3 - nitric acid

b) HNO_2 - nitrous acid

c) H_2SO_4 - sulfuric acid

➢ **Examples:**

ClO^- *hypo*chlor*ite* ion	**HClO *hypo*chlor*ous* acid**
ClO_2^- chlor*ite* ion	**$HClO_2$ chlor*ous* acid**
ClO_3^- chlor*ate* ion	**$HClO_3$ chlor*ic* acid**
ClO_4^- *per*chlor*ate* ion	**$HClO_4$ *per*chlor*ic* acid**

Types of Compounds

Organic compounds contain **carbon** with either H, N, O, or S

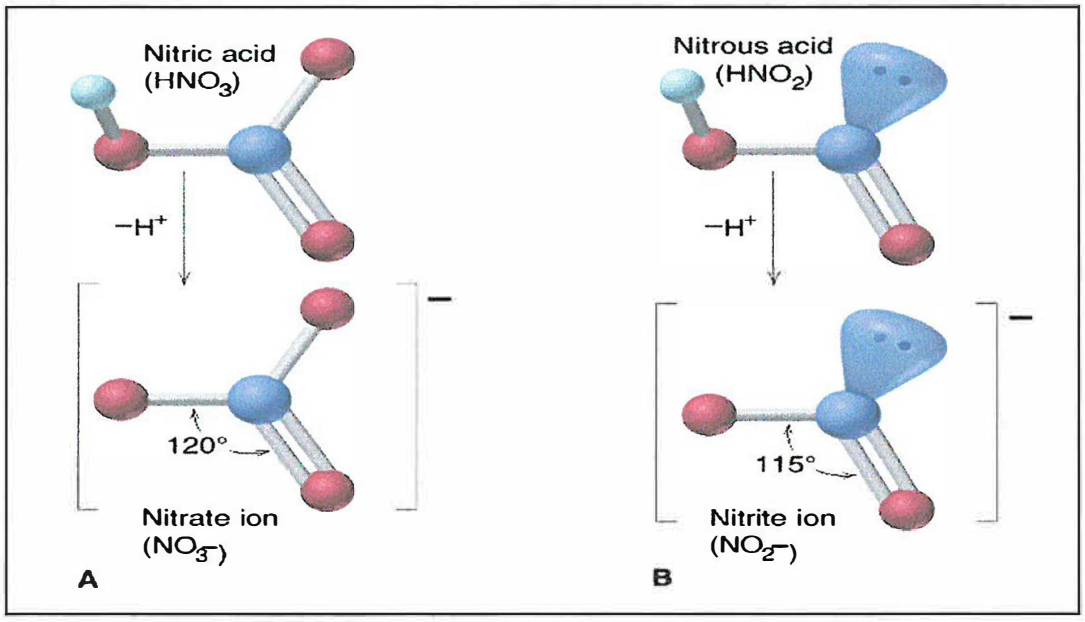
Inorganic compounds are generally placed in 4 classes: Ionic, Molecular, Acids and Bases, & Hydrates.

One class of such a hydroxyl (OH) group acid is the *oxoacids*, in which the acid proton is on a hydroxyl group with an oxo group attached to the same atom. Oxoacids are found where the central atom has a high oxidation number. We may consider two types of oxoacids:

1) Simple oxoacids contain one atom of the parent element. Sulfuric acid, H_2SO_4, is an example of an oxoacid.

Other examples include H_2CO_3, HNO_3, and H_3PO_4. These types of oxoacids are formed by the electronegative element at the upper right of the periodic table and by other elements with high oxidation numbers.

) Substituted oxoacids occur when one or more hydroxyl groups of the oxoacids are replaced by other groups, to give a series of substituted oxoacids. These include fluorosulfuric acid, $O_2SF(OH)$, and aminosulfuric acid, $O_2S(NH_2)OH$.



A *hydrate* is a compound that incorporates water molecules into its fundamental solid structure.

In a hydrate (which usually has a specific crystalline form), a defined number of water molecules are associated with each formula unit of the primary material.

In gypsum, *two* water molecules are present for every formula unit of $CaSO_4$ in the solid; the chemical name of gypsum is calcium sulfate dihydrate: $CaSO_4 \cdot 2H_2O_{(s)}$

<u>Note that the dot, or multiplication sign, indicates that the waters are there.</u>

A few other examples are barium chloride dihydrate - $BaCl_2 \cdot 2H_2O$.
lithium perchlorate trihydrate - $LiClO_4 \cdot 3H_2O$,
and magnesium carbonate pentahydrate - $MgCO_3 \cdot 5H_2O$.

BINARY COVALENT COMPOUNDS

The covalent bond is much harder to break than an ionic bond. The ionic bonds of soluble ionic compounds come apart in water, but covalent bonds do not usually come apart in water. Covalent bonds make real molecules, groups of atoms that are genuinely attached to each other. Binary covalent compounds have two types of atom in them, usually non-metal atoms. Covalent bonds can come in double (sharing of two pairs of electrons) and triple (three pairs of electrons) bonds.

FORMULA	COMMON NAME	SYSTEM NAME
N_2O	nitrous oxide	dinitrogen monoxide
NO	nitric oxide	nitrogen monoxide
N_2O_3	nitrous anhydride	dinitrogen trioxide
NO_2	nitrogen dioxide	nitrogen dioxide
N_2O_4	nitrogen tetroxide	dinitrogen tetroxide
N_2O_5	nitric anhydride	dinitrogen pentoxide
NO_3	nitrogen trioxide	nitrogen trioxide

- **Hydrocarbons: substances composed of only hydrogen and carbon.**

Hydrocarbons usually occur in specific types of rocks, principally shales. Organic remains gradually rot, and are buried and compressed by new sediments. Heat and pressure may change the carbon from the rocks into hydrocarbons. Oil tends to form in rocks whose carbon comes largely from marine plants and animals. Gas tends to form in deposits, which contain carbon largely from land plants.

STUFF I SHOULD KNOW FOR THE AP TEST BUT DO NOT KNOW YET.

IONS LIST

acetate	$C_2H_3O_2^-$	ferric	Fe^{3+} (yellow)	oxalate	$C_2O_4^{2-}$
aluminum	Al^{3+}	ferrous	Fe^{2+} (green)	oxide	O^{2-}
ammonium	NH_4^+	fluoride	F^-	perbromate	BrO_4^-
barium	Ba^{2+}	hydrogen	H^+	perchlorate	ClO_4^-
bicarbonate	HCO_3^-	hydronium	H_3O^+	periodate	IO_4^-
bisulfate	HSO_4^-	hydroxide	OH^-	permanganate	MnO_4^- *(purple)*
bisulfide	HS^-	hypobromite	BrO^-	peroxide	O_2^{2-}
bisulfite	HSO_3^-	hypochlorite	ClO^-	phosphate	PO_4^{3-}
bromate	BrO_3^-	hypoiodite	IO^-	phosphide	P^{3-}
bromide	Br^-	iodate	IO_3^-	phosphite	PO_3^{3-}
bromite	BrO_2^-	iodide	I^-	potassium	K^+
calcium	Ca^{2+}	iodite	IO_2^-	silver	Ag^+
carbonate	CO_3^{2-}	lead	Pb^{2+}	sodium	Na^+
chlorate	ClO_3^-	lithium	Li^+	stannic	Sn^{4+}
chloride	Cl^-	magnesium	Mg^{2+}	stannous	Sn^{2+}
chlorite	ClO_2^-	manganese	Mn^{2+}	strontium	Sr^{2+}
chromate	CrO_4^{2-} (yellow)	mercuric	Hg^{2+}	sulfate	SO_4^{2-}
chromium	Cr^{3+}	mercurous	Hg_2^{2+}	sulfide	S^{2-}
cupric	Cu^{2+} (blue)	nickel	Ni^{2+} (green)	sulfite	SO_3^{2-}
cuprous	Cu^+ (green)	nitrate	NO_3^-	thiocyanate	SCN^-
cyanide	CN^-	nitride	N^{3-}	thiosulfate	$S_2O_3^{2-}$
dichromate	$Cr_2O_7^{2-}$ (orange)	nitrite	NO_2^-	zinc	Zn^{2+}

SOLUBILITY RULES

Always soluble:

alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+), NH_4^+, NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$

Generally soluble: (mnemonics)

Cl^-, Br^-, I^- Soluble except Ag^+, Pb^{2+}, Hg_2^{2+} (AP/H)

F^- Soluble except Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Mg^{2+} (CBS-PM)

SO_4^{2-} Soluble except Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+} (CBS/PBS)

Generally insoluble:

O^{2-}, OH^- Insoluble exceptand alkali metals, and NH_4^+ Ca^{2+}, Sr^{2+}, Ba^{2+} (CBS) somewhat soluble

CO_3^{2-}, PO_4^{3-}, S^{2-}, SO_3^{2-}, $C_2O_4^{2-}$, CrO_4^{2-} Insoluble except alkali metals and NH_4^+

GASES THAT FORM

→ H_2CO_3 → $CO_2 + H_2O$ → NH_4OH → $NH_3 + H_2O$

→ H_2SO_3 → $SO_2 + H_2O$ → H_2S

→ HNO_2 → $NO + NO_2 + H_2O$ → HCN

WEAK ELECTROLYTES

Weak Acids (*esp.* $HC_2H_3O_2$ and HF)

(Memorize the 8 strong acids... all others are weak)

HCl	hydrochloric acid	HNO_3	nitric acid
HBr	hydrobromic acid	HIO_4	periodic acid
HI	hydroiodic acid	H_2SO_4	sulfuric acid
$HClO_4$	perchloric acid	$HClO_3$	chloric acid

Ammonium Hydroxide ($NH_4OH \approx NH_3(aq)$) Water (H_2O)

DRIVING FORCES -- Double Replacement

- Insoluble Solid (Precipitate)
- Weak Electrolyte (H_2O or Weak Acid)
- Gas Formation

STRONG OXIDIZERS (Oxidizing Agents)

MnO_4^- in acid solution	→ $Mn^{2+} + H_2O$
MnO_2 in acid solution	→ $Mn^{2+} + H_2O$
MnO_4^- in neutral or basic sol'n	→ MnO_2
$Cr_2O_7^{2-}$ in acid solution	→ $Cr^{3+} + H_2O$
$Cr_2O_7^{2-}$ with a base	→ $CrO_4^{2-} + H_2O$
CrO_4^{2-} in basic solution	→ $CrO_2^- + H_2O$
HNO_3, concentrated	→ $NO_2 + H_2O$
HNO_3, dilute (e.g. 6 M)	→ $NO + H_2O$
H_2SO_4, hot, concentrated	→ $SO_2 + H_2O$
Free halogens (e.g. Cl_2)	→ halide ions (Cl^-)
H_2O_2 in acid solution	→ H_2O
Note: H_2O_2 decomposes	→ $H_2O + O_2$
Na_2O_2	→ NaOH
$HClO_4$	→ $Cl^- + H_2O$

Other Oxidizers

Metal-"ic" ions (e.g. Sn^{4+}, Fe^{3+})	→ "-ous" ions (Sn^{2+}, Fe^{2+})
H_2O	→ $H_2 + OH^-$

STRONG REDUCERS (Reducing Agents)

Halide ions (e.g. Cl^-)	→ Free halogen (Cl_2)
Free metals	→ metal ions
"ites" SO_3^{2-} or SO_2, NO_2^-	→ "ates" SO_4^{2-}, NO_3^-
Free halogens, dil. basic sol'n	→ hypohalite ions (ClO^-)
Free halogens, conc. basic sol'n	→ halate ions (ClO_3^-)
$S_2O_3^{2-}$	→ $S_4O_6^{2-}$

Other Reducers

Metal-"ous" ions (e.g. Sn^{2+})	→ "-ic" ions (Sn^{4+})
H_2O	→ $O_2 + H^+$

Complex Ions & Common Ligands

Ligands	polar molecules & anions	NH_3, H_2O, OH^-, CN^-, Cl^-	**Odd example:** $Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$
Central Ions	transition metals and Al^{3+}	Ag^+, Cu^{2+}, Ni^{2+}, Zn^{2+}, etc. & Al^{3+}	
Examples	usually twice the number of ligands as the charge on the central ion. **Key Words:** "excess, concentrated"	$Ag(CN)_2^-$, $Cu(NH_3)_4^{2+}$, $Ni(OH)_4^{2-}$, $Zn(NH_3)_4^{2+}$, $Al(OH)_6^{3-}$	**Reaction with Acid:** $Cu(NH_3)_4^{2+} + H^+ \rightarrow Cu^{2+} + NH_4^+$

Organic Chemistry & Functional Groups

alkanes	alkenes	alkynes	aromatics (benzene)
C_nH_{2n+2}	C_nH_{2n}	C_nH_{2n-2}	C_6H_6
alcohol	**aldehyde**	**ketone**	**ether**
R — OH	R — C(=O) — H	R — C(=O) — R	R — O — R
carboxylic acid	**ester**	**amine**	**amide**
R — C(=O) — OH	R — C(=O) — O — R	R — NH_2	R — C(=O) — NH_2
Substituted benzene:	**ortho = 1,2**	**meta = 1,3**	**para = 1,4**

nuclear chem

alpha $^{4}_{2}He$

beta/electron $^{0}_{-1}e$

neutron $^{1}_{0}n$

positron $^{0}_{+1}e$

ΔH	ΔS	Spont.?
–	+	at all temps
+	+	high temps
–	–	low temps
+	–	no temps

Note: ΔS in J
ΔG & ΔH in kJ

K_{sp} & Solubility, s

1:1 $K_{sp} = s^2$
1:2 $K_{sp} = 4s^3$
1:3 $K_{sp} = 27s^4$
2:3 $K_{sp} = 108s^5$

Lewis Acids & Bases

$BF_3 + NH_3 \rightarrow BF_3NH_3$

acid anhydrides (oxides of nonmetals, CO_2)

basic anhydrides (oxides of metals, MgO)

$MgO + CO_2 \rightarrow MgCO_3$

decomposition reactions: $MgCO_3 \rightarrow MgO + CO_2$

Strange Examples: $P_4O_{10} + H_2O \rightarrow H_3PO_4$

Strange Ions: (nitride, N^{3-}) (hydride, H^-)

$Li + N_2 \rightarrow Li_3N$ $LiH + H_2O \rightarrow H_2 + Li^+ + OH^-$

Electrochemical Cells

anode	cathode
oxidation	reduction
– side	+ side
lower E°	higher E°
e^- leave	e^- enter

Bond Orders

bond	B.O.	
single	1	σ
double	2	σ, π
triple	3	σ, π, π

Flame Test Colors

Barium – green
Sodium – yellow
Copper – blue (w/ green)
Potassium – lavender
Strontium – red
Lithium – red
Calcium – orange

Quantum Numbers

n	1, 2, 3, …
l	0 … (n–1)
m_l	–l … +l
m_s	+½, –½

l	0 = s, 1 = p, 2 = d, 3 = f

SN & hybridization & shape

Steric Number	hybridization	basic shape
1	s	—
2	sp	linear
3	sp^2	△ planar
4	sp^3	tetrahedral
5	sp^3d	△ bipyramidal
6	sp^3d^2	octahedral

Writing Lewis Structures

hint: use one valence electron to connect F's or Cl's then determine lone pairs (Ex: XeF_4)

Product–Favored (Spontaneous) Reactions

$\Delta G < 0$ $E° > 0$ $K_{eq} > 1$

IMF's

London	nonpolar molecules, ex: CH_4, He
dipole–dipole	polar molecules, ex: H_2S, SO_2
hydrogen bonding	H–F, H–O–, H–N–, NH_3, H_2O amines and alcohols
metallic	metals, Ag, Pb
ionic	salts, NaCl, $CaCO_3$ (Note: "ates" contain covalent bonds)
covalent network	C(graphite), C(diamond), SiO_2, WC, Si, SiC (Note: graphite = London, too)

Properties Indicate Strength of Intermolecular Forces (IMF's)

IMF	BP	FP	H_{vap}	H_{fus}	VP
IMF	BP	FP	H_{vap}	H_{fus}	VP

Activity of Metals (Four Groups)

Metals	React with...
Groups I & II	H_2O ex: $Li + H_2O \rightarrow Li^+ + OH^- + H$
all others	Non-oxidizing Acid, ex: HCl $Zn + 2HCl \rightarrow H_2 + ZnCl_2$
Cu, Ag, Hg	Oxidizing Acid, HNO_3 or H_2SO_4 (conc.) $Cu + HNO_3 \rightarrow NO_2 + H_2O + Cu^{2+}$
Au, Pt, Ir	Aqua Regia (HNO_3 + HCl)

Orders of Reactions & Graphs That Give Straight Lines

0 Order	1st Order	2nd Order
[R] vs. Time	ln[R] vs. Time	1/[R] vs. Time
slope = -k	slope = -k	slope = k

Laboratory Glassware and Other Apparatus

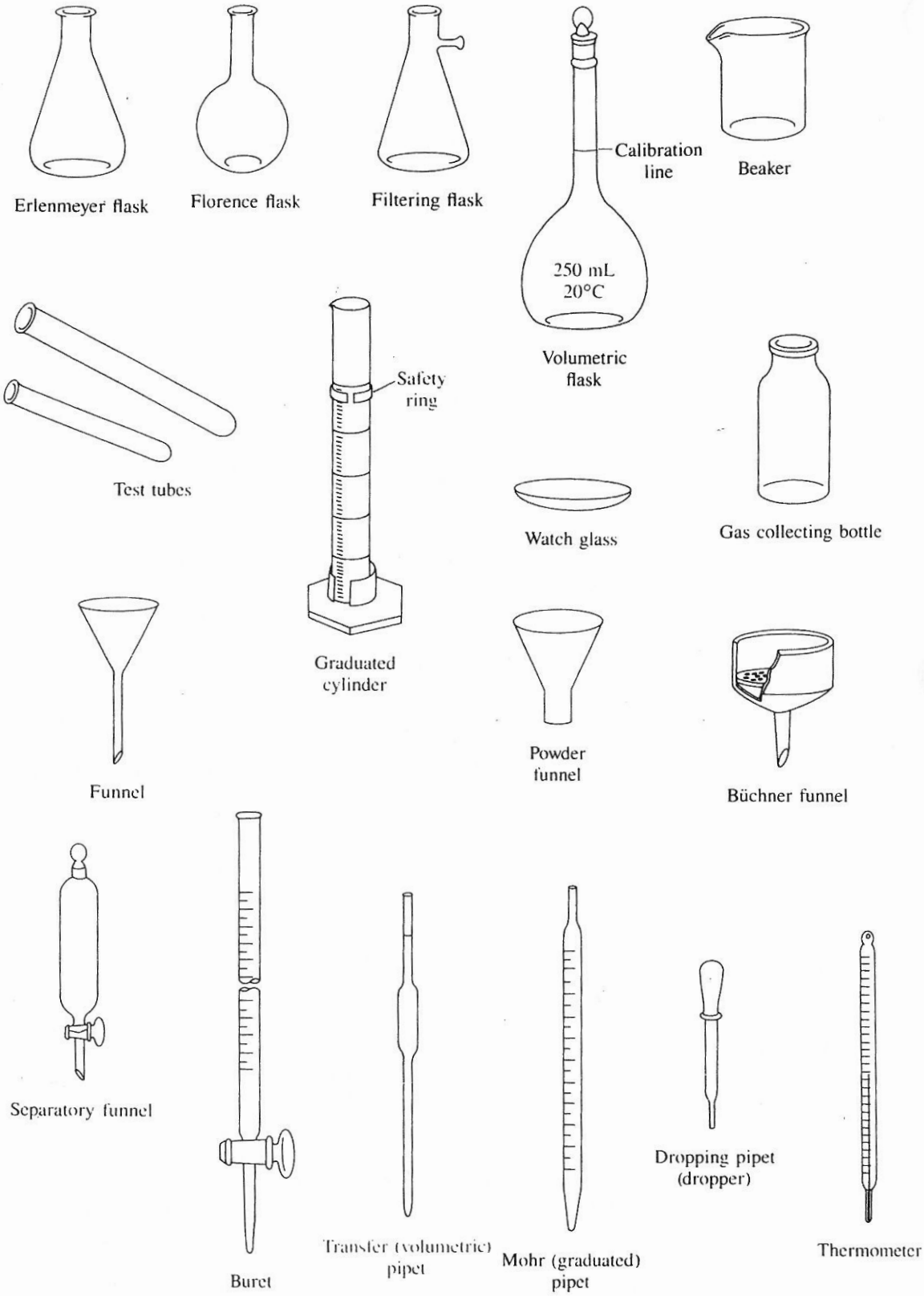
Introduction

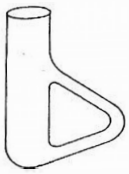
When you open your laboratory locker for the first time, you are likely to be confronted with a bewildering array of various sizes and shapes of glassware and other apparatus. Glass is used more than any other material for the manufacture of laboratory apparatus because it is relatively cheap, usually easy to clean (and keep clean), and can be heated to fairly high temperatures or cooled to quite low temperatures. Most importantly, glass is used because it is impervious to and nonreactive with most reagents encountered in the beginning chemistry laboratory. Most glassware used in chemistry laboratories is made of borosilicate glass, which is relatively sturdy and safe to use at most temperatures. Such glass is sold under such trade names as Pyrex® (Corning Glass Co.) and Kimax® (Kimble Glass Co.). If any of the glassware in your locker does not have either of these trade names marked on it, consult with your instructor before using it at temperatures above room temperature.

Many pieces of laboratory glassware and apparatus have special names that have evolved over the centuries chemistry has been studied. You should learn the names of the most common pieces of laboratory apparatus to make certain that you use the correct equipment for the experiments in this manual. On the next few pages are drawings of common pieces of laboratory apparatus. Compare the equipment in your locker with these drawings, and identify all the pieces of apparatus with which you have been provided. It might be helpful for you to label the equipment in your locker, at least for the first few weeks of the term.

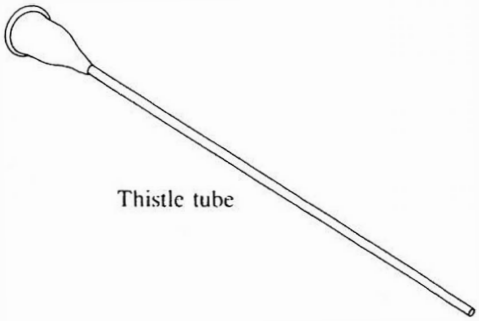
While your locker may not contain all the apparatus in the drawings, be sure to ask your instructor for help if there is equipment in your locker that is not described in the drawings. While examining your locker equipment, watch for chipped, cracked, or otherwise imperfect glassware. *Imperfect glassware is a major safety hazard, and must be discarded.* Don't think that because a beaker has just a little crack, it can still be used. *Replace all glassware that has any cracks, chips, star fractures, or any other deformity.* Most college and university laboratories will replace imperfect glassware free of charge during the first laboratory meeting of the term, but may assess charges if breakage is discovered later in the term.

You may wish to clean the set of glassware in your locker during the first meeting of your laboratory section. This is certainly admirable, but be warned that lab glassware has an annoying habit of becoming dirty again without apparent human intervention. Glassware always looks clean when wet, but tends to dry with water spots (and show minor imperfections in the glass that are not visible when wet). It is recommended that you thoroughly clean all the locker glassware at the start, and then rinse the glassware out before its first use. In the future, clean glassware before leaving lab for the day, and rinse before using. Instructions for the proper cleaning of laboratory glassware are provided after the diagrams of apparatus.

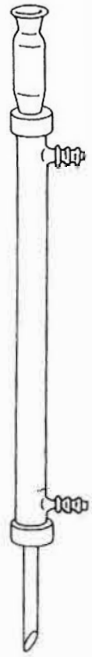

Erlenmeyer flask
Florence flask
Filtering flask
Calibration line
250 mL
20°C
Volumetric flask
Beaker
Test tubes
Safety ring
Graduated cylinder
Watch glass
Gas collecting bottle
Funnel
Powder funnel
Büchner funnel
Separatory funnel
Buret
Transfer (volumetric) pipet
Mohr (graduated) pipet
Dropping pipet (dropper)
Thermometer



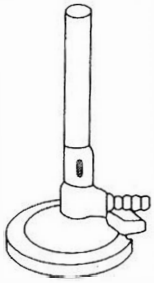
Thiele tube



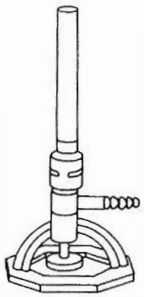
Thistle tube



Liebig condenser
(water-cooled)

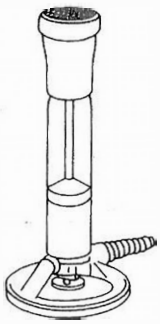


Bunsen
burner

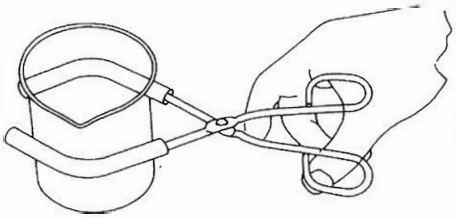


Tirrill burner

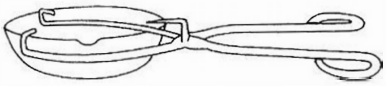
Wing top
(flame spreader)



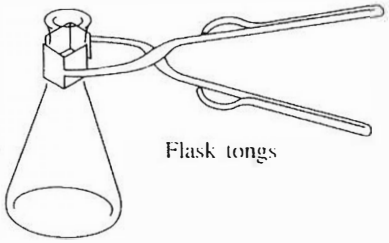
Meker burner



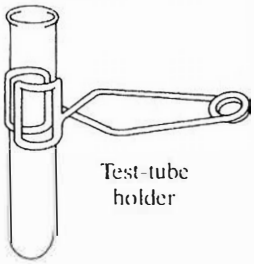
Beaker tongs



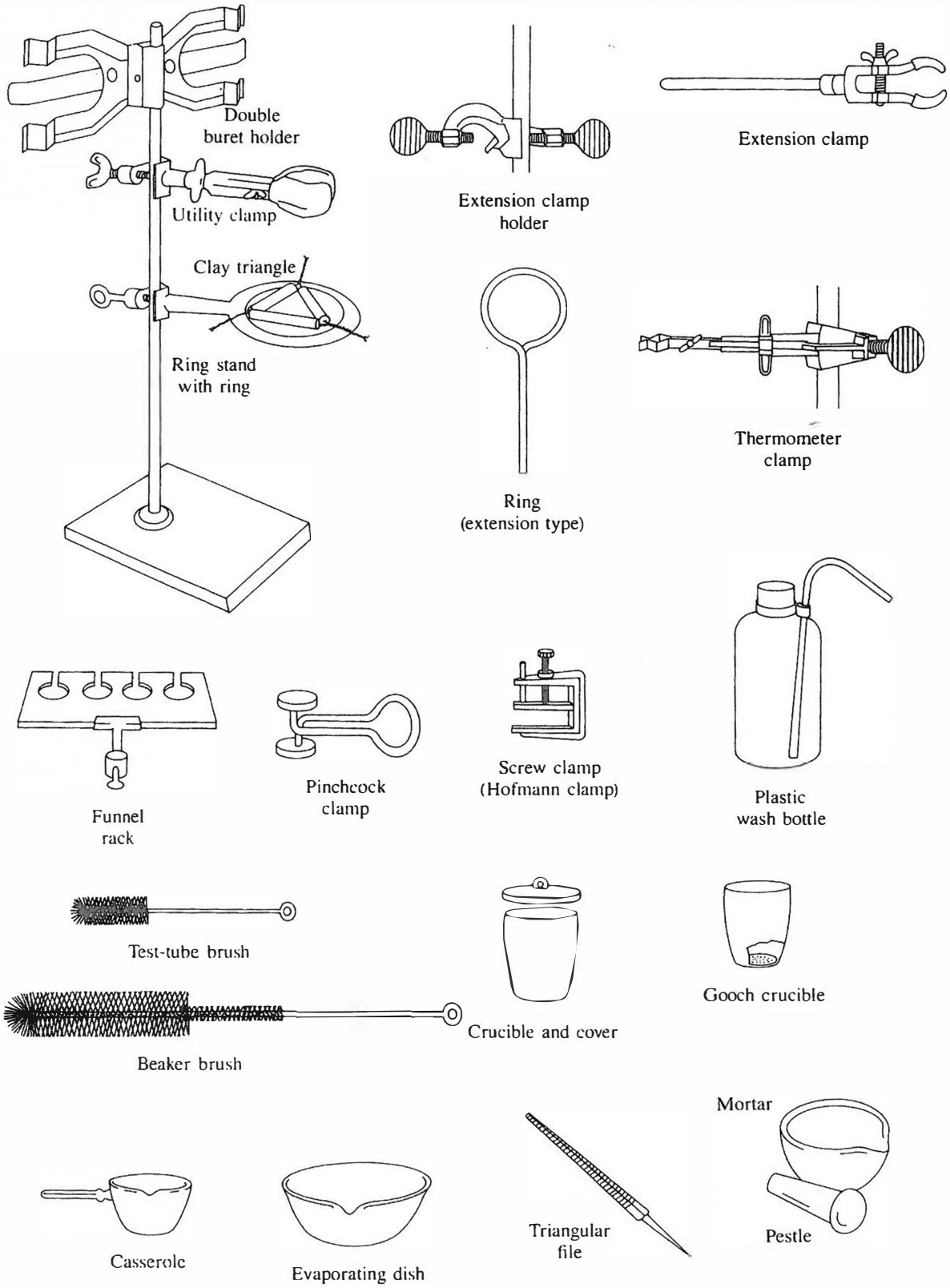
Dish tongs



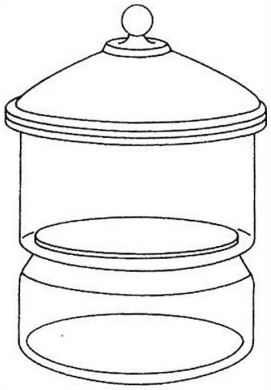
Flask tongs



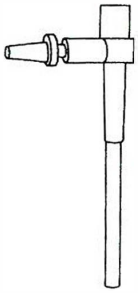
Test-tube
holder


Double
buret holder
Utility clamp
Clay triangle
Ring stand
with ring
Extension clamp
holder
Extension clamp
Ring
(extension type)
Thermometer
clamp
Funnel
rack
Pinchcock
clamp
Screw clamp
(Hofmann clamp)
Plastic
wash bottle
Test-tube brush
Beaker brush
Crucible and cover
Gooch crucible
Casserole
Evaporating dish
Triangular
file
Mortar
Pestle

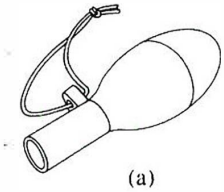
Forceps



Desiccator

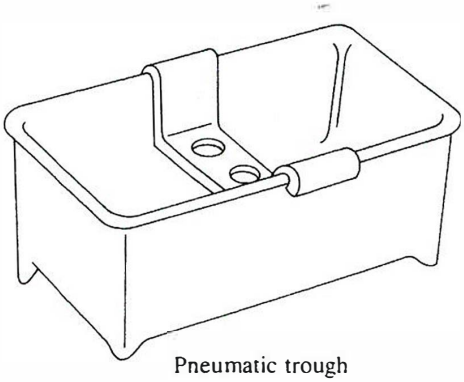


Water aspirator
(filter pump)

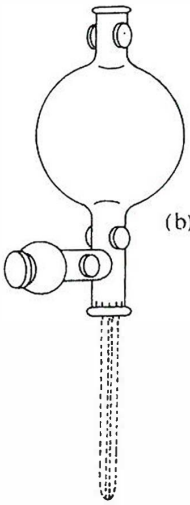


(a)

Rubber
policeman

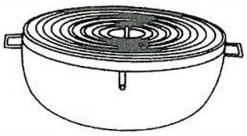


Pneumatic trough

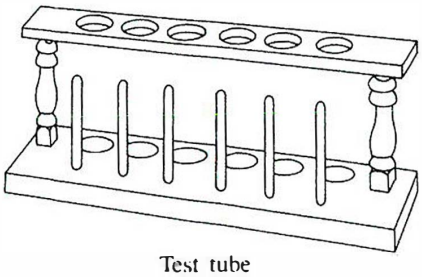


(b)

Pipet
safety bulbs

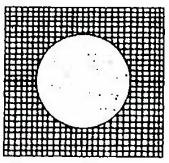


Water bath



Test tube
rack

Spatulas



Wire gauze
(ceramic center)

2019 Richard Montgomery HS
AP Chemistry Summer Assignment

Chapter 1

1.11 Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: (a) rice pudding, (b) seawater, (c) magnesium, (d) gasoline).

1.15 Give the chemical symbol or name for the following elements, as appropriate: (a) carbon, (b) nitrogen, (c) bromine, (d) zinc, (e) iron, (f) P, (g) Ca, (h) He, (i) Pb, (j) Ag.

1.19 Label each of the following as either a physical process or a chemical process: (a) corrosion of aluminum metal, (b) melting of ice, (c) pulverizing an aspirin, (d) digestion of a candy bar, (e) explosion of nitroglycerin.

1.25 Make the following conversions: (a) 62°F to °C, (b) 216.7°C to °F, (c) 233°C to K, (d) 315 K to °F, (e) 2500°F to K.

1.35 What is the number of significant figures in each of the following measured quantities? (a) 358kg, (b) 0.054 s, (c) 6.3050 cm, (d) 0.0105 L, (e) 7.0500×10^{-3} m^3.

1.36 Indicate the number of significant figures in each of the following measured quantities: (a) 3.774 km, (b) 205 m^2, (c) 1.700 cm, (d) 350.00 K, (e) 307.080 g.

1.37 Round each of the following numbers to four significant figures, and express the result in standard exponential notation: (a) 102.53070, (b) 656,980, (c) 0.008543210, (d) 0.000257870, (e) -0.0357202.

1.39 Carry out the following operations, and express the answers with the appropriate number of significant figures.
(a) 12.0550 + 9.05
(b) 257.2 – 19.789
(c) $(6.21 \times 10^3)(0.1050)$
(d) 0.0577/0.753

1.46 Carry out the following conversions: (a) 0.105 in. to mm, (b) 0.650 qt to mL, (c) 8.75 μm/s to km/hr, (d) 1.955 m^3 to yd^3, (e) $3.99/lb to $/kg, (f) 8.75 lb/ft^3 to g/mL.

1.72 Automobile batteries contain sulfuric acid, which is commonly referred to as battery acid. Calculate the number of grams of sulfuric acid in .500 L of battery acid if the solution has a density of 1.28 g/mL and is 38.1% sulfuric acid by mass.

Chapter 2

2.22 (a) Which two types of the following are isotopes of the same element ${}^{31}_{16}X$, ${}^{31}_{15}X$, ${}^{32}_{16}X$? (b) What is the identity of the element whose isotopes you have selected?

2.31 Only two isotopes of copper occur naturally, ${}^{63}Cu$ (atomic mass = 62.9296 amu; abundance 69.17%) and ${}^{65}Cu$ (atomic mass = 64.9278 amu; abundance 30.83%). Calculate the atomic weight (average atomic mass) of copper.

2.41 What can we tell about a compound when we know the empirical formula? What additional information is conveyed by the molecular formula? By the structural formula? Explain in each case.

2.59 Predict whether each of the following compounds is molecular or ionic: (a) B_2H_6, (b) CH_3OH, (c) $LiNO_3$, (d) Sc_2O_3, (e) CsBr, (f) NOCl (g) NF_3, (h) Ag_2SO_4.

2.66 Name the following ionic compounds: (a) K_2O, (b) $NaClO_2$, (c) $Sr(CN)_2$, (d) $CoOH_2$, (e) $Fe_2(CO_3)_3$, (f) $Cr(NO_3)_3$, (g) $(NH_4)_2SO_3$, (h) NaH_2PO_4, (i) $KMnO_4$, (j) $Ag_2Cr_2O_7$.

2.68 Give the chemical formula for each of the following ionic compounds: (a) sodium phosphate, (b) potassium sulfate, (c) copper(I) oxide, (d) zinc nitrate, (e) mercury(II) bromide, (f) iron(III) carbonate, (g) sodium hypobromite.

2.87 Identify the element represented by each of the following symbols and give the number of protons and neutrons in each: (a) ${}^{74}_{33}X$, (b) ${}^{127}_{53}X$, (c) ${}^{152}_{63}X$, (d) ${}^{209}_{83}X$.

2.90 The element (Pb) consists of four naturally occurring isotopes with atomic masses 203.97302, 205.97444, 206.97587, 207.97663 amu. The relative abundances of these four isotopes are 1.4, 24.1, 22.1 and 52.4%, respectively. From these data, calculate the atomic weight of lead.

Chapter 3

3.23 Calculate the percentage of mass of oxygen in the following compounds: (a) morphine, $C_{17}H_{19}NO_3$; (b) codeine, $C_{18}H_{21}NO_3$; (c) cocaine, $C_{17}H_{21}NO_4$; (d) tetracycline, $C_{22}H_{24}N_2O_8$.

3.35 (a) What is the mass, in grams, of 2.50×10^{-3} mol of ammonium phosphate? (b) How many moles of chloride ions are in 0.2550 g of aluminum chloride? (c) What is the molar mass, in grams, of 7.70×10^{20} molecules of caffeine, $C_8H_{10}N_4O_2$? (d) What is the molar mass of cholesterol if 0.00105 mol weighs 0.406 g?

3.42 At least 25 μg of tetrahydrocannabinol (THC), the active ingredient in marijuana, is required to produce intoxication. The molecular formula of THC is $C_{21}H_{30}O_2$;. How many moles of THC does 25 μg represent? How many molecules?

3.46 Determine the empirical formulas of the compounds with the following compositions by mass:
(a) 55.3% K, 14.6% P, 30.1% O
(b) 24.5% Na, 14.9% Si, and 60.6% F
(c) 62.1% C, 5.21% H, 12.1% N, and 20.7% O

3.49 Determine the empirical and molecular formulas of each of the following substances:
(a) Ibuprofen, a headache remedy, contains 75.69% C, 8.80% H, and 15.51% O by mass, and has a molar mass of 206 g/mol.
(b) Cadavarine, a foul smelling substance produced by the action of bacteria on meat, contains 58.55% C, 13.81% H, and 27.40% N by mass; its molar mass is 102.2 g/mol.
(c) Epinephrine (adrenaline), a hormone secreted into the bloodstream in times of danger or stress, contains 59.0% C, 7.1% H, 26.2% O, and 7.7% N by mass; its MW is about 180 amu.

3.59 Several brands of antacids use $Al(OH)_3$ to react with stomach acid, which contains primarily HCl:

$$Al(OH)_3\,(s) + HCl\,(aq) \rightarrow AlCl_3\,(aq) + H_2O\,(l)$$

(a) Balance the equation.
(b) Calculate the number of grams of HCl that can react with .500 g of $Al(OH)_3$.
(c) Calculate the number of grams of $AlCl_3$ and the number of grams of H_2O formed when .500 g of $Al(OH)_3$ reacts.
(d) Show your calculations in parts (b) and (c) are consistent with the law of conservation of mass.

3.71 Sodium hydroxide reacts with carbon dioxide as follows:

$$2NaOH\,(s) + CO_2\,(g) \rightarrow Na_2CO_3\,(s) + H_2O\,(l)$$

Which reagent is the limiting reactant when 1.85 mol NaOH and 1.00 mol CO_2 are allowed to react? How many moles of Na_2CO_3 can be produced? How many moles of the excess reactant remain after the completion of the reaction?

3.73 The fizz produced when an Alka-Seltzer tablet is dissolved in water is due to the reaction between sodium bicarbonate and citric acid:

$$3NaHCO_3\,(aq) + H_3C_6H_5O_7\,(aq) \rightarrow 3CO_2\,(g) + 3H_2O\,(l) + Na_3C_6H_5O_7\,(aq)$$

In a certain experiment 1.00 g of sodium bicarbonate and 1.00 g of citric acid are allowed to react. (a) Which is the limiting reactant? (b) How many grams of carbon dioxide form? (c) How many grams of the excess reactant remain after the limiting reactant is completely consumed?

3.88 Vanillin, the dominant flavoring in vanilla, contains C, H, & O. When 1.05 g of this substance is completely combusted, 2.43 g of CO_2 and 0.50 g of H_2O are produced. What is the empirical formula of vanillin?

3.91 An element X forms an iodide (XI_3) and a chloride (XCl_3). The iodide is quantitatively converted to the chloride when it is heated in a stream of chlorine.

$$2XI_3\,(aq) + 3Cl_2\,(aq) \rightarrow 3CO_2\,(g) + 2XCl_3 + 3I_2$$

If 0.5000 g of XI_3 is treated, 0.2360 g of XCl_3 is obtained. (a) Calculate the atomic weight of the element X. (b) Identify the element X.

3.100 Aspirin ($C_9H_8O_4$) is produced from salicylic acid ($C_7H_6O_3$) and acetic anhydride ($C_4H_6O_3$):

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + HC_2H_3O_2$$

(a) How much salicylic acid is required to produce 150 kg of aspirin, assuming that all of the salicylic acid is converted to aspirin? (b) How much salicylic acid would be required if only 80% of the salicylic acid is converted to aspirin? (c) What is the theoretical yield of aspirin if 185 kg of salicylic acid is allowed to react with 125 kg of acetic anhydride? (d) If the situation described in part (c) produces 182 kg of aspirin, what is the percentage yield?

Richard Montgomery HS: Lab Safety Rules

1. Students will work individually on some experiments, while other experiments will be done with partners or groups. For partner or group work each student should interpret data and answer questions separately unless directed otherwise.

2. Be prepared to work when you arrive at the laboratory. Familiarize yourself with the lab procedures before beginning the lab.

3. Carefully follow directions, both written and oral. Do only the steps described in the procedure of the experiment or that are described and/or approved by the teacher. If you are in doubt about any procedure, ask your teacher for help.

4. Everyone should be alert and proceed with caution at all times in the laboratory. Take care not to bump another student and to remain at your lab station while performing an experiment. An unattended experiment can result in an accident.

5. Wear safety glasses/goggles whenever you are in the lab. Aprons are required for some experiments. You may wear aprons at your discretion for any experiment to avoid staining or ruining your clothing.

6. Clothing should be appropriate for working in the lab. Jackets, ties, and other loose garments should be removed. Ideally, dress for lab should include long pants and shoes which cover the entire foot.

7. Do not engage in horseplay such as tickling, throwing objects, squirting water, etc. Some of the most serious lab accidents have resulted due to this type of behavior. Misbehavior such as horseplay could result in your dismissal from the classroom and you may not be allowed to participate in future labs and activities.

8. If you cut yourself, spill a chemical on yourself, or receive a burn by touching a hot object, run cold water over the affected area, and you or your partner notify your teacher immediately.

9. Learn the location of the eye wash fountain and/or water faucets in your school lab. If a substance is splashed in your eyes, immediately use the eye wash fountain or a water faucet to rinse your eyes. This should not happen if you are wearing goggles. Wear safety glasses/goggles at all times in the lab unless you are specifically told by your instructor that you do not need to wear them.

10. **Do not taste, touch, or smell any reagents** unless directed to do so by your teacher. When smelling chemicals or gases, use a wafting motion to direct the odor toward your nose.

11. Extreme caution should be used when using a Bunsen burner. Keep your head and clothing away from the flame and turn off the burner when it is not in use. Long hair should be tied back to avoid it catching on fire. If your clothing should catch fire, stop, drop, and roll while your lab partner notifies the instructor. Before leaving the lab, check to see that all gas valves and hot plates are turned off.

12. Keep flammable and combustible materials away from open flames. Some examples of flammable materials include alcohol, carbon disulfide, and acetone.

13. When heating a substance in a test tube, be careful not to point the mouth of the test tube at another person or yourself.

14. Use caution and the proper equipment to handle hot objects. Cool glass looks just the same as hot glass.

15. Handle chemicals carefully. Check the label of all bottles before removing the contents. Take only as much as your need. Do not return unused chemicals to reagent bottles. Report all spills or incorrect procedures to the teacher.

16. Handle toxic or combustible gases or chemicals only under the direction of the teacher. Use the fume hood when using these materials or when directed to do so.

17. Know the correct procedure for mixing acid solutions. **ALWAYS add the acid slowly to the water**. Never add water to a large amount of acid.

18. Never handle broken glass with your bare hands. Use a brush and dustpan to clean up broken glass. Dispose of the glass as directed by your teacher. Record and report all breakage or loss of apparatus to your teacher.

19. Breakage fees for glassware and equipment broken due to misuse will be charged to the student(s) responsible for the broken item(s) at the replacement cost.

20. Keep insoluble waste material out of the sink. Dispose of waste material as instructed by your teacher.

21. Work areas should be clean and tidy at all times. Only lab procedures, lab notebooks, pencils and and sometimes calculators should be brought to the work area.

22. Aseptic technique (hand washing with antibacterial soap before and after the lab, disinfection of tables before and after the lab, and using the proper procedures for handling microbes) should be followed at all times when handling microbes) should be followed at all times when handling bacteria, protozoans, and fungi. Notify your teacher before you begin the lab of any health problems you have which may have compromised your immune system.

23. When an experiment is completed, always clean equipment and return it to the proper place. Clean your lab table.

24. Wash hands thoroughly with soap and water before leaving the lab.

25. Drinking during the lab is prohibited. Talking quietly is permitted, but it should not interfere with your work.

26. Do not let the potential hazards listed above make you afraid to participate in the lab. If instructions are followed and care is taken, the likelihood of an accident is greatly reduced. Labs are usually the most fun-filled part of any science course.