AP Chemistry
Summer 2018 Assignment

Welcome to AP Chemistry! I am very excited to have you in my class this year; but in order to succeed you need to understand my expectations of you. This class is going to be regarded as a college level class. I will be preparing you for the AP Chemistry Exam, which receiving a high enough grade on will earn you college credit. If you are not in this class with the goal of learning at a college level and receiving college credit, I would reassess your decision. In this class you will be expected to do college level work, it is not going to be like the Chemistry class you just finished. You will be expected to be at a higher academic and maturity level than the rest of your class because I know you are the best at SCS. We have a lot of material to cover this year, and in order to best prepare you for the AP test; work must be done on time and you will need to take responsibility for your learning and prepare yourself outside of the class.

You need to start preparing for this class this summer. I will expect you to have a solid background of Chemistry the day you walk into this class. In order to do this, I have put together this summer assignment.

Due Date:
The assignment will be due the first day of school. There will also be a test on the concepts found in this packet within the first week of school.
There will be no late work accepted, both for this assignment and the rest of the year in this class. You should be prepared enough that if any issues arise that they can be handled before the due date. I do not want to hear any excuses.

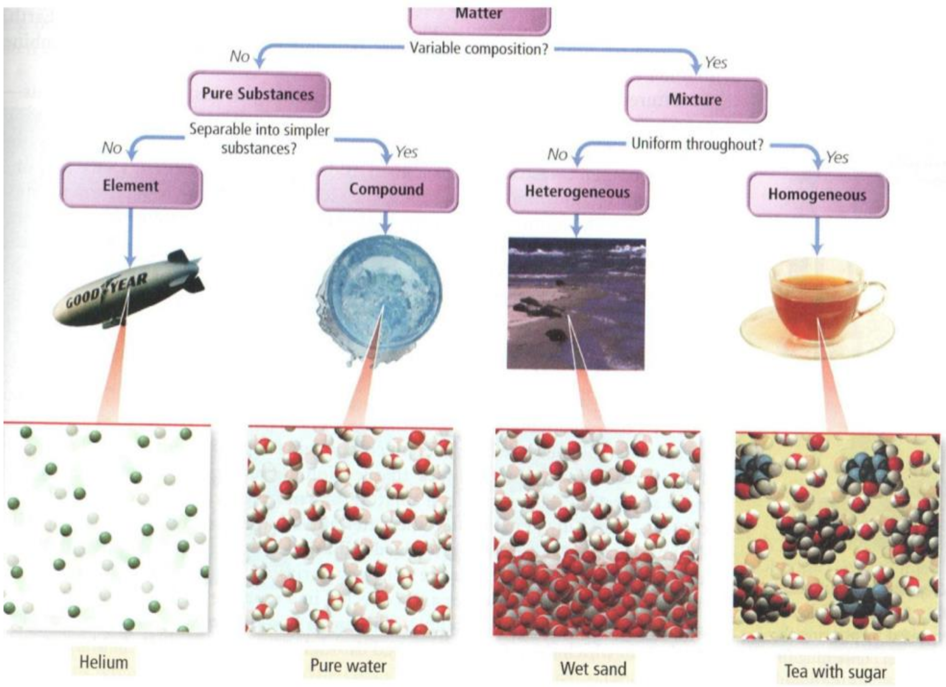
Assignment:
- All the concepts in this packet I require you to know by the first day of class. Anything within this packet is testable material.
- The problems should be completed with ALL work shown on an attached separate sheet of paper. Please keep all work neat and legible. Circle all answers. Be prepared to do similar problems on a test.
- At the back of this packet, there is a blank Periodic Table, complete it. Include, the element symbol (and name if you can fit it and it still looks neat), atomic mass, and atomic number. Also color-code the metals, metalloids, nonmetals, and liquids. Label the following groups: halogens, alkali metals, transition metals, alkaline earth metals, and noble gases.

Question:
Please feel free to email me if you have any questions. My email is dtakanishi@southlandscs.com.

Concepts:

1. Know the differences between the 3 states of matter. (Solid, liquid, gas)

2. Know the following chart:



3. Difference between a physical and chemical change, find an example of each
   Physical change: alters only the state or appearance, not composition
   Chemical change: alters the composition of matter

4. Know the following definitions, and be able to distinguish between the different types of energies if examples were given.
   Energy: capacity to do work
   Work: action or force through a distance (Ex: push a box across the floor)
   Kinetic Energy: energy associated with motion (Ex: a ball rolling)
   Potential Energy: energy associated with its position or composition (Ex: a weight being help above the ground)
   Thermal Energy: energy associated with the temperature of an object

5. Give examples in chemistry of how these energies can be converted to one another. Note: The law of conservation of energy.

6. Know how to convert from Celsius to Kelvin. Know the boiling and freezing temperatures of water on both scales.

7. Know the following chart:

TABLE 1.2 SI Prefix Multipliers

Prefix	Symbol	Multiplier	
exa	E	1,000,000,000,000,000,000	(10^{18})
peta	P	1,000,000,000,000,000	(10^{15})
tera	T	1,000,000,000,000	(10^{12})
giga	G	1,000,000,000	(10^{9})
mega	M	1,000,000	(10^{6})
kilo	k	1,000	(10^{3})
deci	d	0.1	(10^{-1})
centi	c	0.01	(10^{-2})
milli	m	0.001	(10^{-3})
micro	μ	0.000001	(10^{-6})
nano	n	0.000000001	(10^{-9})
pico	p	0.000000000001	(10^{-12})
femto	f	0.000000000000001	(10^{-15})
atto	a	0.000000000000000001	(10^{-18})

8. Density = mass/volume. Density of water is 1g/1mL.

9. Memorize this conversion factor: 1mL = 1cm^3
   Other conversion factors, do not need to memorize, but will need in the problems:
   - 1 kilometer (km) = 0.6214 mile (mi)
   - 1 meter (m) = 39.37 inches (in) = 1.094 yards (yd)
   - 1 foot (ft) = 30.48 centimeters (cm)
   - 1 inch (in) = 2.54 centimeters (cm)
   - 1 kilogram (kg) = 2.205 pounds (lb)
   - 1 pound (lb) = 453.59 grams (g)
   - 1 ounce (oz) = 28.35 grams (g)
   - 1 liter (L) = 1000 mL = 1000 cm^3
   - 1 liter (L) = 1.057 quarts (qt)
   - 1 U.S. gallon (gal) = 3.785 liters (L)

10. Define and differentiate between accuracy and precision.
    Accuracy: how close the measured value is to the actual value.
    Precise: how close a series of measurements are to one another or how reproducible they are.

11. Use DIMENSIONAL ANALYSIS when converting units. (This is very important!) The following pages will give a good overview of dimensional analysis as well as example problems.

The unit, in, cancels and we are left with cm as our final unit. The quantity $\frac{2.54\ \text{cm}}{1\ \text{in}}$ is a **conversion factor**—a fractional quantity with the units we are *converting from* on the bottom and the units we are *converting to* on the top. Conversion factors are constructed from any two equivalent quantities. In this example, 2.54 cm = 1 in, so we construct the conversion factor by dividing both sides of the equality by 1 in and canceling the units

$$2.54\ \text{cm} = 1\ \text{in}$$

$$\frac{2.54\ \text{cm}}{1\ \text{in}} = \frac{1\ \cancel{\text{in}}}{1\ \cancel{\text{in}}}$$

$$\frac{2.54\ \text{cm}}{1\ \text{in}} = 1$$

The quantity $\frac{2.54\ \text{cm}}{1\ \text{in}}$ is equivalent to 1, so multiplying by the conversion factor affects only the units, not the actual quantity. To convert the other way, from centimeters to inches, we must—using units as a guide—use a different form of the conversion factor. If you accidentally use the same form, you will get the wrong result, indicated by erroneous units. For example, suppose that you want to convert 31.8 cm to inches.

$$31.8\ \text{cm} \times \frac{2.54\ \text{cm}}{1\ \text{in}} = \frac{80.8\ \text{cm}^2}{\text{in}}$$

The units in the above answer (cm^2/in), as well as the value of the answer, are obviously wrong. When you solve a problem, always look at the final units. Are they the desired units? Always look at the magnitude of the numerical answer as well. Does it make sense? In this case, our mistake was the form of the conversion factor. It should have been inverted so that the units cancel as follows:

$$31.8\ \cancel{\text{cm}} \times \frac{1\ \text{in}}{2.54\ \cancel{\text{cm}}} = 12.5\ \text{in}$$

Conversion factors can be inverted because they are equal to 1 and the inverse of 1 is 1. Therefore,

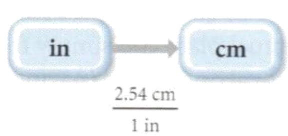
$$\frac{2.54\ \text{cm}}{1\ \text{in}} = 1 = \frac{1\ \text{in}}{2.54\ \text{cm}}$$

Most unit conversion problems take the following form:

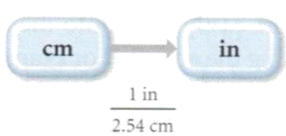
$$\text{Information given} \times \text{conversion factor(s)} = \text{information sought}$$

$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

In this book, we diagram a problem solution using a *conceptual plan*. A conceptual plan is a visual outline that helps you to see the general flow of the problem. For unit conversions, the conceptual plan focuses on units and the conversion from one unit to another. The conceptual plan for converting in to cm is as follows:



The conceptual plan for converting the other way, from cm to in, is just the reverse, with the reciprocal conversion factor:



Each arrow in a conceptual plan for a unit conversion has an associated conversion factor with the units of the previous step in the denominator and the units of the following step in the numerator. In the following section, we incorporate the idea of a conceptual plan into an overall approach to solving numerical chemical problems.

Procedure for Solving Unit Conversion Problems

Sort

Begin by sorting the information in the problem into *given* and *find*.

Strategize

Devise a *conceptual plan* for the problem. Begin with the *given* quantity and symbolize each conversion step with an arrow. Below each arrow, write the appropriate conversion factor for that step. Focus on the units. The conceptual plan should end at the *find* quantity and its units. In these examples, the other information needed consists of relationships between the various units as shown.

Solve

Follow the conceptual plan. Begin with the *given* quantity and its units. Multiply by the appropriate conversion factor(s), canceling units, to arrive at the *find* quantity.

Round the answer to the correct number of significant figures by following the rules in Section 1.7. Remember that exact conversion factors do not limit significant figures.

Check

Check your answer. Are the units correct? Does the answer make physical sense?

EXAMPLE 1.7 Unit Conversion

Convert 1.76 yards to centimeters.

Given 1.76 yd

Find cm

Conceptual Plan

$$\text{yd} \xrightarrow[\frac{1\text{ m}}{1.094\text{ yd}}]{} \text{m} \xrightarrow[\frac{100\text{ cm}}{1\text{ m}}]{} \text{cm}$$

Relationships Used

1.094 yd = 1 m
1 m = 100 cm
(These conversion factors are from Tables 1.2 and 1.3.)

Solution

$$1.76\ \cancel{\text{yd}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{100\text{ cm}}{1\ \cancel{\text{m}}} = 160.8775\text{ cm}$$

$$160.8775\text{ cm} = 161\text{ cm}$$

The units (cm) are correct. The magnitude of the answer (161) makes physical sense because a centimeter is a much smaller unit than a yard.

For Practice 1.7

Convert 288 cm to yards.

EXAMPLE 1.8 Unit Conversion

Convert 1.8 quarts to cubic centimeters.

Given 1.8 qt

Find cm^3

Conceptual Plan

$$\text{qt} \xrightarrow[\frac{1\text{ L}}{1.057\text{ qt}}]{} \text{L} \xrightarrow[\frac{1000\text{ mL}}{1\text{ L}}]{} \text{mL} \xrightarrow[\frac{1\text{ cm}^3}{1\text{ mL}}]{} \text{cm}^3$$

Relationships Used

1.057 qt = 1 L
1 L = 1000 mL
1 ml = 1 cm^3
(These conversion factors are from Tables 1.2 and 1.3.)

Solution

$$1.8\ \cancel{\text{qt}} \times \frac{1\ \cancel{\text{L}}}{1.057\ \cancel{\text{qt}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \frac{1\text{ cm}^3}{1\ \cancel{\text{mL}}} = 1.70293 \times 10^3\text{ cm}^3$$

$$1.70293 \times 10^3\text{ cm}^3 = 1.7 \times 10^3\text{ cm}^3$$

The units (cm^3) are correct. The magnitude of the answer (1700) makes physical sense because a cubic centimeter is a much smaller unit than a quart.

For Practice 1.8

Convert 9255 cm^3 to gallons.

Units Raised to a Power

When building conversion factors for units raised to a power, remember to raise both the number and the unit to the power. For example, to convert from in^2 to cm^2, we construct the conversion factor as follows:

$$\begin{aligned} 2.54\text{ cm} &= 1\text{ in} \\ (2.54\text{ cm})^2 &= (1\text{ in})^2 \\ (2.54)^2\text{ cm}^2 &= 1^2\text{ in}^2 \\ 6.45\text{ cm}^2 &= 1\text{ in}^2 \\ \frac{6.45\text{ cm}^2}{1\text{ in}^2} &= 1 \end{aligned}$$

The following example shows how to use conversion factors involving units raised to a power.

EXAMPLE 1.9 Unit Conversions Involving Units Raised to a Power

Calculate the displacement (the total volume of the cylinders through which the pistons move) of a 5.70-L automobile engine in cubic inches.

Sort Sort the information in the problem into *Given* and *Find*.

Given 5.70 L

Find in^3

Strategize Write a conceptual plan. Begin with the given information and devise a path to the information that you are asked to find. Notice that for cubic units, the conversion factors must be cubed.

Conceptual Plan

$$\text{L} \xrightarrow{\frac{1\ \text{mL}}{10^{-3}\ \text{L}}} \text{mL} \xrightarrow{\frac{1\ \text{cm}^3}{1\ \text{mL}}} \text{cm}^3 \xrightarrow{\frac{(1\ \text{in})^3}{(2.54\ \text{cm})^3}} \text{in}^3$$

Relationships Used

$1\ \text{mL} = 10^{-3}\ \text{L}$

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

$2.54\ \text{cm} = 1\ \text{in}$

(These conversion factors are from Tables 1.2 and 1.3.)

Solve Follow the conceptual plan to solve the problem. Round the answer to three significant figures to reflect the three significant figures in the least precisely known quantity (5.70 L). These conversion factors are all exact and therefore do not limit the number of significant figures.

Solution

$$5.70\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mL}}}{10^{-3}\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times \frac{(1\ \text{in})^3}{(2.54\ \cancel{\text{cm}})^3} = 347.835\ \text{in}^3$$

$$= 348\ \text{in}^3$$

Check The units of the answer are correct and the magnitude makes sense. The unit cubic inches is smaller than liters, so the volume in cubic inches should be larger than the volume in liters.

For Practice 1.9

How many cubic centimeters are there in 2.11 yd^3?

For More Practice 1.9

A vineyard has 145 acres of Chardonnay grapes. A particular soil supplement requires 5.50 grams for every square meter of vineyard. How many kilograms of the soil supplement are required for the entire vineyard? ($1\ \text{km}^2 = 247$ acres)

12. Memorize these polyatomic ions, know how to go from the name to the compound (including the charge) and compound to the name:
    - Sulfate (SO_4^{2-})
    - Sulfite (SO_3^{2-})
    - Nitrate (NO_3^{-})
    - Nitrite (NO_2^{-})
    - Ammonium (NH_4^{+})
    - Perchlorate (ClO_4^{-})
    - Chlorate (ClO_3^{-})
    - Chlorite (ClO_2^{-})
    - Hydroxide (OH^{-})
    - Phosphate (PO_4^{3-})
    - Carbonate (CO_3^{2-})

Hydrogen Carbonate (Bicarbonate) (HCO_3^-)
Cyanide (CN^-)
Acetate ($C_2H_3O_2^-$)
Chromate (CrO_4^{2-})
Permanganate (MnO_4^-)

13. Know the diatomic molecules. O_2, N_2, H_2, F_2, Cl_2, Br_2, and I_2

14. Understand the subatomic particles: protons, neutrons, and electrons.
    a. Know their charges
    b. Know their masses
    c. What significance do they play in regards to an atom? Proton defines what element it is, neutron creates isotopes, and electrons create ions.

15. Mass number is calculated by adding the number of protons and neutrons within an atom. (Be able to calculate).

16. Know the following ways isotopes are written:

For neon, with 10 protons, the mass numbers of the three different naturally occurring isotopes are 20, 21, and 22, corresponding to 10, 11, and 12 neutrons, respectively.

Isotopes are often symbolized in the following way:

Mass number → $^{A}_{Z}X$ ← Chemical symbol
Atomic number →

where X is the chemical symbol, A is the mass number, and Z is the atomic number. Therefore, the symbols for the neon isotopes are

$$^{20}_{10}Ne \quad ^{21}_{10}Ne \quad ^{22}_{10}Ne$$

Notice that the chemical symbol, Ne, and the atomic number, 10, are redundant: if the atomic number is 10, the symbol must be Ne. The mass numbers, however, are different for different isotopes, reflecting the different number of neutrons in each one.

A second common notation for isotopes is the chemical symbol (or chemical name) followed by a dash and the mass number of the isotope.

Chemical symbol or name → X-A ← Mass number

In this notation, the neon isotopes are

Ne-20	Ne-21	Ne-22
neon-20	neon-21	neon-22

We can summarize what we have learned about the neon isotopes in the following table:

Symbol	Number of Protons	Number of Neutrons	A (Mass Number)	Natural Abundance(%)
Ne-20 or $^{20}_{10}Ne$	10	10	20	90.48
Ne-21 or $^{21}_{10}Ne$	10	11	21	0.27
Ne-22 or $^{22}_{10}Ne$	10	12	22	9.25

Notice that all isotopes of a given element have the same number of protons (otherwise they would be different elements). Notice also that the mass number is the *sum* of the number of protons and the number of neutrons. The number of neutrons in an isotope is the difference between the mass number and the atomic number ($A - Z$). The different isotopes of an element generally exhibit the same chemical behavior—the three isotopes of neon, for example, all exhibit the same chemical inertness.

17. Losing and gaining electrons creates ions. Cations are positively charged, anions are negatively charged.

18. Atomic mass is the average mass of an element. Be able to calculate the atomic mass, given the isotope masses and their abundance.

An important part of Dalton's atomic theory was that all atoms of a given element have the same mass. However, in Section 2.6, we learned that because of isotopes, the atoms of a given element often have different masses, so Dalton was not completely correct. We can, however, calculate an average mass—called the **atomic mass**—for each element.

The atomic mass of each element is listed directly beneath the element's symbol in the periodic table and represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope*. For example, the periodic table lists the atomic mass of chlorine as 35.45 amu. Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). Its atomic mass is computed as follows:

17
Cl
35.45
chlorine

$$\text{Atomic mass} = 0.7577(34.97\text{ amu}) + 0.2423(36.97\text{ amu}) = 35.45\text{ amu}$$

Notice that the atomic mass of chlorine is closer to 35 than 37. Naturally occurring chlorine contains more chlorine-35 atoms than chlorine-37 atoms, so the weighted average mass of chlorine is closer to 35 amu than to 37 amu.

19. Understand the mole and Avogadro's number. Be able to use Avogadro's number in conversions.

The Mole: A Chemist's "Dozen"

When we count large numbers of objects, we often use units such as a dozen (12 objects) or a gross (144 objects) to organize our counting and to keep our numbers smaller. With atoms, quadrillions of which may be in a speck of dust, we need a much larger number for this purpose. The chemist's "dozen" is called the **mole** (abbreviated mol) and is defined as the *amount* of material containing 6.0221421×10^{23} particles.

$$1\text{ mol} = 6.0221421 \times 10^{23}\text{ particles}$$

This number is also called **Avogadro's number**, named after Italian physicist Amedeo Avogadro (1776–1856), and is a convenient number to use when working with atoms, molecules, and ions. In this book, we will usually round Avogadro's number to four significant figures or 6.022×10^{23}. Notice that the definition of the mole is an *amount* of a substance. We will often refer to the number of moles of substance as the *amount* of the substance.

The first thing to understand about the mole is that it can specify Avogadro's number of anything. For example, 1 mol of marbles corresponds to 6.022×10^{23} marbles, and 1 mol of sand grains corresponds to 6.022×10^{23} sand grains. *One mole of anything is 6.022×10^{23} units of that thing.* One mole of atoms, ions, or molecules, however, makes up objects of everyday sizes. For example, 22 copper pennies contain approximately 1 mol of copper atoms and a tablespoon of water contains approximately 1 mol of water molecules.

The second, and more fundamental, thing to understand about the mole is how it gets its specific value.

> **The numerical value of the mole is defined as being equal to the number of atoms in exactly 12 grams of pure carbon-12 (12 g C = 1 mol C atoms = 6.022×10^{23} C atoms).**

The definition of the mole gives us a relationship between mass (grams of carbon) and number of atoms (Avogadro's number). This relationship, as we will see shortly, allows us to count atoms by weighing them.

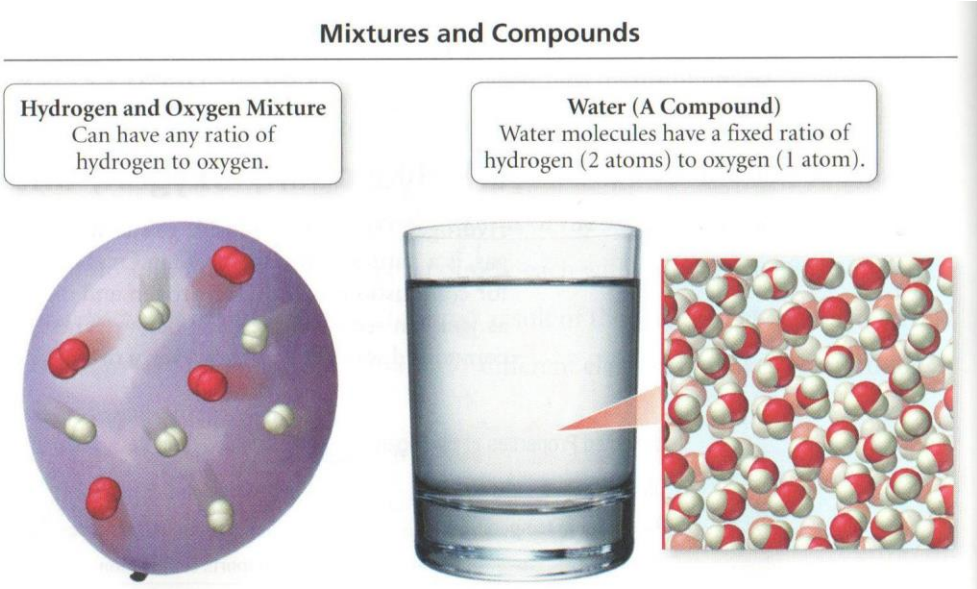
Twenty-two copper pennies contain approximately 1 mol of copper atoms.

Beginning in 1982, pennies became almost all zinc, with only a copper coating. Before 1982, however, pennies were mostly copper.

One tablespoon of water contains approximately one mole of water molecules.

One tablespoon is approximately 15 mL; one mole of water occupies 18 mL.

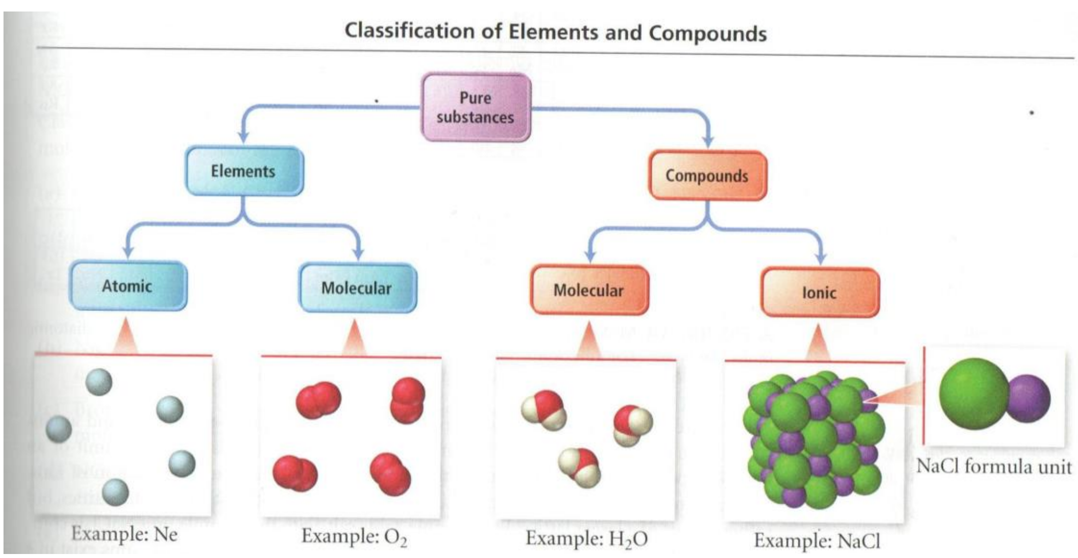
20. Difference between a mixture and compound:



21. **Ionic Bond**: oppositely charged ions are attracted to one another by electrostatic forces (cation and anion)
    Covalent Bond: electron(s) are shared between two atoms

22. **Empirical Formula**: relative number of atoms of each element in a molecule (lowest common denominator)
    Molecular Formula: actual number of atoms of each element in a molecule
    Example: Empirical formula for hydrogen peroxide is HO, but the molecular formula is H_2O_2

23. Understand the following chart:
    Atomic elements: single atom
    Molecular elements: when an element exits naturally as a molecule (diatomic molecules)
    Molecular compounds: two or more covalently bonded nonmetals
    Ionic compound: cations and anions bound with ionic bonds



Problems:

1. Classify each of the following as a pure substance or a mixture. If it is a pure substance, classify it as an element or a compound. If it is a mixture, classify it as homogeneous or heterogeneous.

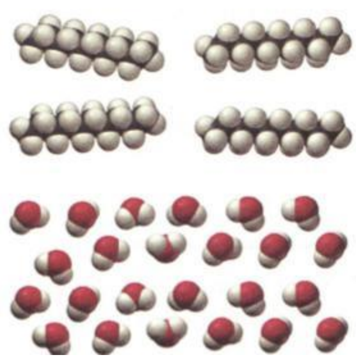
   a. sweat **b.** carbon dioxide
   c. aluminum **d.** vegetable soup

2. Complete the following table.

Substance	Pure or mixture	Type (element or compound)
aluminum	pure	element
apple juice	___________	___________
hydrogen peroxide	___________	___________
chicken soup	___________	___________

3. Classify each of the following molecular diagrams as a pure substance or a mixture. If it is a pure substance, classify it as an element or a compound. If it is a mixture, classify it as homogeneous or heterogeneous.


(a)


(b)

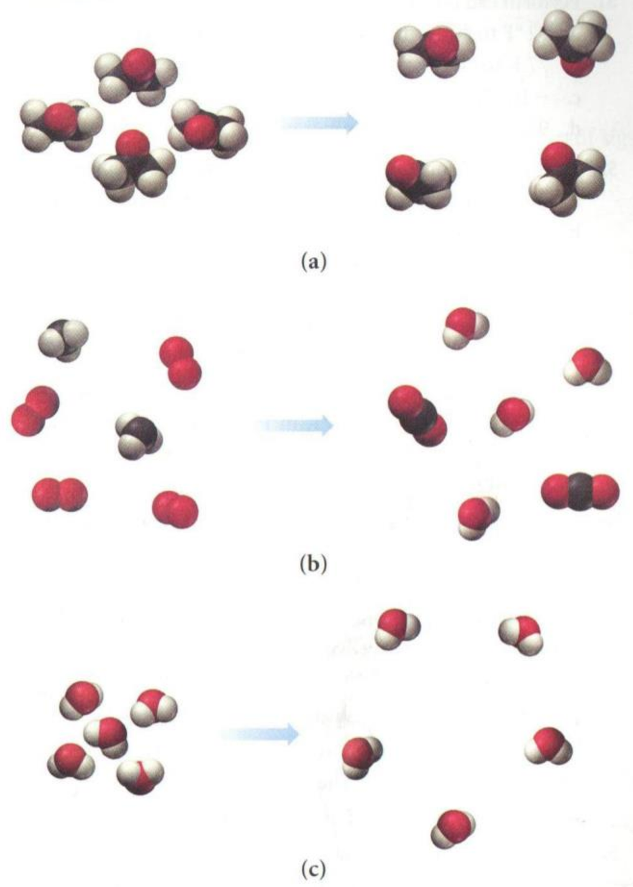

(c)


(d)

4. Several properties of isopropyl alcohol are listed below. Classify each of the properties as physical or chemical.
    a. Colorless
    b. Flammable
    c. Liquid at room temperature
    d. Density = 0.79 g/mL
    e. Mixes with water

5. Classify each of the following changes as physical or chemical.
    a. Natural gas burns in a stove
    b. The liquid propane in a gas grill evaporates because the user left the valve open
    c. The liquid propane in a gas grill burns in a flame
    d. A bicycle frame rusts on repeated exposure to air and water.

6. 49. Based on the molecular diagram, classify each change as physical or chemical.



7. Complete the following table:

a.	1245 kg	1.245×10^6 g	1.245×10^9 mg
b.	515 km	_________dm	__________cm
c.	122.355 s	_________ms	__________ks
d.	3.345 kJ	_________J	__________mJ

8. Glycerol is a syrupy liquid often used in cosmetics and soaps. A 3.25 L sample of pure glycerol has a mass of 4.10×10^3 g. What is the density of glycerol in g/cm^3?

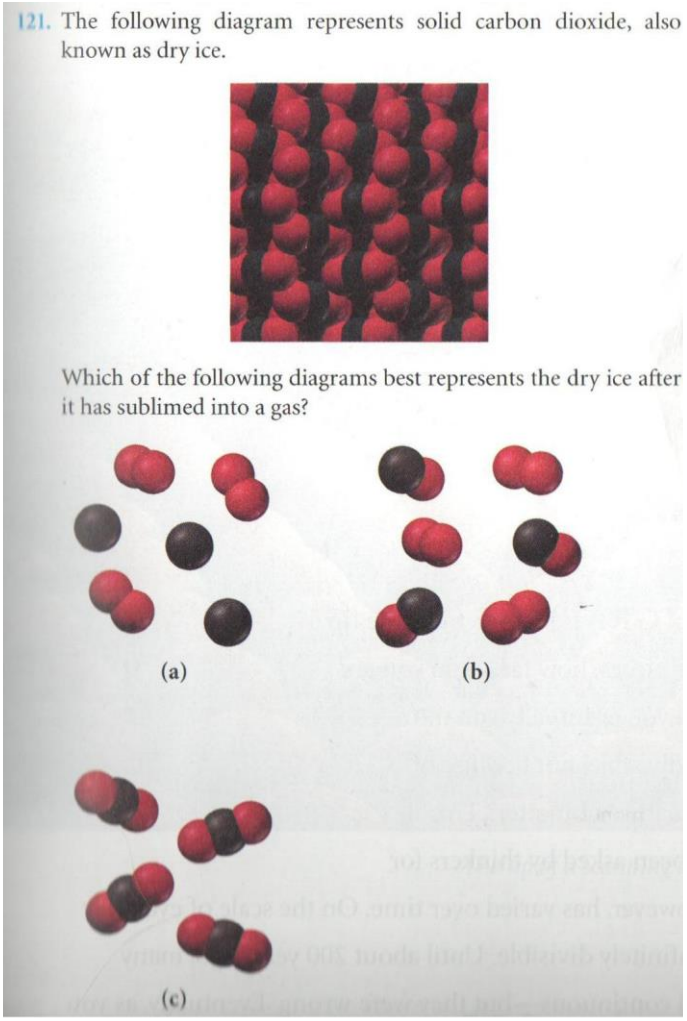
9. Ethylene glycol has a density of 1.11 g/cm^3.
   a. What is the mass in g of a 417 mL of this liquid?
   b. What is the volume in L of 4.1 kg of this liquid?

10. Perform each of the following conversions:
    a. 154 cm to in.
    b. 3.14 kg to g
    c. 3.5 L to qt
    d. 109 mm to in
11. A modest-sized house has an area of 195 m^2. What is its area in:
    a. km^2
    b. dm^2
    c. cm^2

12. There are 60 seconds in a minute, 60 minutes in an hour, 24 hours in a solar day, and 365.24 solar days in a solar year. How many seconds are in a solar year?

13. The density of titanium is 4.51 g/cm^3. What is the volume (in cubic inches) of 3.5 lb of titanium?

14.



121. The following diagram represents solid carbon dioxide, also known as dry ice.

Which of the following diagrams best represents the dry ice after it has sublimed into a gas?

(a)

(b)

(c)

15. Which of the following statements are true?
    a. If an atom has an equal number of protons and electrons, the atom is charge-neutral.
    b. Electrons are attracted to protons.
    c. Electrons are much lighter than neutrons.
    d. Protons have twice the mass of neutrons.

16. Write isotopic symbols of the form $^{A}_{Z}X$ for each of the following isotopes.
    a. The sodium isotope with 12 neutrons
    b. The oxygen isotope with 8 neutrons
    c. The aluminum isotope with 14 neutrons
    d. The iodine isotope with 74 neutrons

17. Determine the number of protons and neutrons in each of the following isotopes.
    a. $^{14}_{7}N$ b. $^{23}_{11}Na$ c. $^{222}_{86}Rn$ d. $^{208}_{82}Pb$

18. Determine the number of protons and electrons in each of the following ions.
    a. Ni^{2+}
    b. S^{2-}
    c. Br^{-}
    d. Cr^{3+}
19. Predict the charge of the ion formed by each of the following elements.
    a. O
    b. K
    c. Al
    d. Rb
20. Fill in the Blanks to complete the following table:

Symbol	Ion Formed	Number of Electrons in Ion	Number of Protons in Ion
Ca	Ca^{2+}	______	______
______	Be^{2+}	2	______
Se	______	______	34
In	______	______	49
Cl	______	______	17
Te	______	54	______
Br	Br^{-}	______	______
______	Sr^{2+}	______	38

21. Which of the following pairs of elements do you expect to be most similar? Why?
    a. N and Ni
    b. Mo and Sn
    c. Na and Mg
    d. Cl and F
    e. Si and P

22. Rubidium has two naturally occurring isotopes with the following masses and natural abundances:
    Rb-85, mass of 84.9118,
    72.15% Rb-87, mass of 86.9092,
    27.85% Calculate the atomic mass of
    rubidium.

23. How many sulfur atoms are there in 3.8 mol of sulfur?

24. What is the amount in moles of each of the following?
    a. 11.8 g Ar
    b. 3.55 g Zn
    c. 26.1 g Ta

d. 0.211 g Li

25. What is the mass in grams of each of the following?
    a. 2.3×10^{-3} mol Sb
    b. 0.0355 mol Ba
    c. 43.9 mol Xe
    d. 1.3 mol W

26. How many silver atoms are in 3.78 g of silver?

27. How many carbon atoms are there in a diamond with a mass of 52 mg?

28. Fill in the blanks of the following table:

Symbol	Z	A	Number of Protons	Number of Electrons	Number of Neutrons	Charge
______	8	______	______	______	8	2-
Ca^{2+}	20	______	______	______	20	______
Mg^{2+}	______	25	______	______	13	2+
N^{3-}	______	24	______	10	______	______

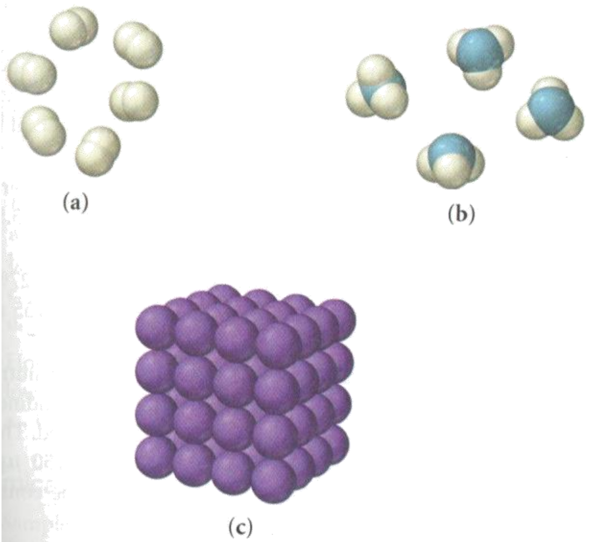
29. Without doing calculations, determine which of the following samples contains the greatest amount of the element in moles.
    a. 55.0 g Cr
    b. 45.0 g Ti
    c. 60.0 g Zn

30. Determine the number of each type of atom in each of the following formulas:
    a. $Ca_3(PO_4)_2$
    b. $SrCl_2$
    c. KNO_3
    d. $Mg(NO_2)_2$

31. Classify each of the following elements as atomic or molecular.
    a. Neon
    b. Fluoride
    c. Potassium
    d. Nitrogen

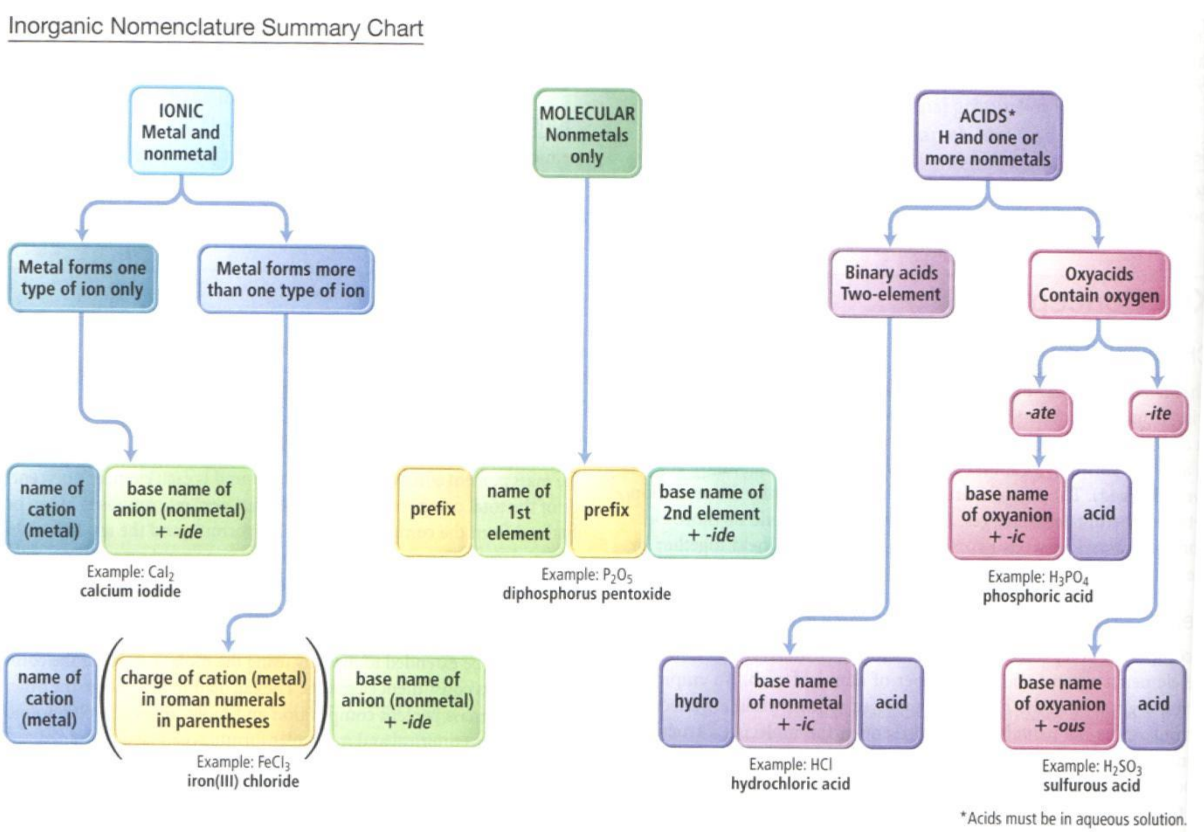
32. Classify each of the following compounds as ionic or molecular.
    a. CO_2
    b. $NiCl_2$
    c. NaI
    d. PCl_3

33. Based on the following molecular views, classify each substance as an atomic element, a molecular element, an ionic compound, or a molecular compound.



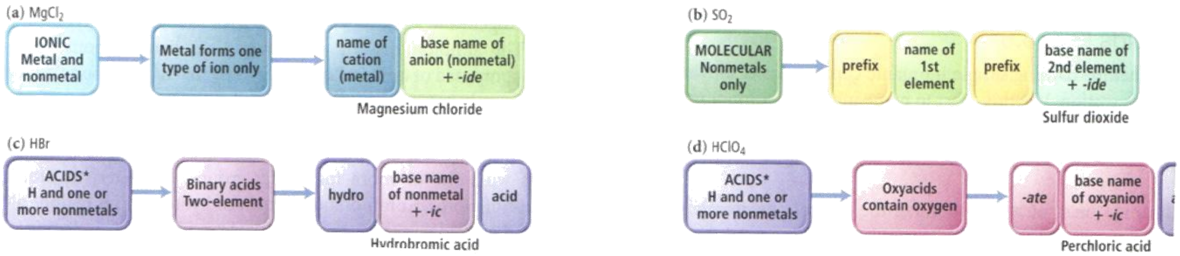
Naming and Balancing Chemical Equations:

These topics are very important, and it will help a lot that you start practicing throughout the summer. Unfortunately, the AP books are not in yet, so you can't read the section on it. But I am assuming you have done some of this in your previous chemistry class. Use the flowchart below to help you work through the problems (try everyone!). This will not be on the test until I go over it with you and I believe you fully understand it.



Using the Flowchart

The examples below show how to name compounds using the flowchart. The path through the flowchart is shown below each compound followed by the correct name for the compound.



1. Write the formula for the ionic compound that forms between each of the following pairs of elements.
   a. Magnesium and sulfur
   b. Barium and oxygen
   c. Strontium and bromine
   d. Beryllium and chloride

2. Write a formula for the compound that forms between barium and each of the polyatomic ions.
   a. Hydroxide
   b. Chromate
   c. Phosphate
   d. Cyanide

3. Name each of the following ionic compounds.
   a. Mg_3N_2
   b. KF
   c. Na_2O
   d. Li_2S
   e. $SnCl_4$
   f. PbI_2
   g. Fe_2O_3
   h. CuI_2
   i. SnO
   j. Cr_2S_3
   k. RbI
   l. $BaBr_2$
   m. $CuNO_2$
   n. $Mg(C_2H_3O_2)_2$
   o. $Ba(NO_3)_2$
   p. $Pb(C_2H_3O_2)_2$
   q. $KClO_3$
   r. $PbSO_4$

4. Write a formula for each of the following ionic compounds:
    a. Sodium hydrogen sulfite
    b. Lithium permanganate
    c. Silver nitrate
    d. Potassium sulfate
    e. Rubidium hydrogen sulfate
    f. Potassium hydrogen carbonate

5. Name each of the following molecular compounds.
    a. CO
    b. NI_3
    c. $SiCl_4$
    d. N_4Se_4
    e. I_2O_5

6. Write a formula for each of the following molecular compounds
    a. Phosphorous trichlorate
    b. Chlorine monoxide
    c. Disulfur tetrafluoride
    d. Phosphorous pentafluoride
    e. Diphosphorous pentasulfide

7. Name each of the following acids.
    a. HI
    b. HNO_3
    c. H_2CO_3
    d. $HC_2H_3O_2$

8. Write formulas for each of the following acids.
    a. Hydrofluoric acid
    b. Hydrobromic acid
    c. Sulfurous acid

9. Balance each of the following chemical equations
    a. $CO_{2(g)} + CaSiO_{3(s)} + H_2O_{(l)}$ ✐ $SiO_{2(s)} + Ca(HCO_3)_{2(aq)}$
    b. $Co(NO_3)_{3(aq)} + (NH_4)_2S_{(aq)}$ ✐ $Co_2S_{3(s)} + NH_4NO_{3(aq)}$
    c. $Cu_2O_{(s)} + C_{(s)}$ ✐ $Cu_{(s)} + CO_{(g)}$
    d. $H_{2(g)} + Cl_{2(g)}$ ✐ $HCl_{(g)}$

10. Balance each of the following chemical equations
    a. $Na_2S_{(aq)} + Cu(NO_3)_{2(aq)}$ ✐ $NaNO_{3(aq)} + CuS_{(s)}$
    b. $N_2H_{4(l)}$ ✐ $NH_{3(g)} + N_{2(g)}$
    c. $HCl_{(aq)} + O_{2(g)}$ ✐ $H_2O_{(l)} + Cl_{2(g)}$
    d. $FeS_{(s)} + HCl_{(aq)}$ ✐ $FeCl_{2(aq)} + H_2S_{(g)}$

11. Write a balanced chemical equation for each of the following:
    a. Solid lead(II) sulfide reacts with aqueous hydrobromic acid to form solid lead(II) bromide and dihydrogen monosulfide gas.
    b. Gaseous carbon monoxide reacts with hydrogen gas to form gaseous methane (CH_4) and liquid water.
    c. Aqueous hydrochloric acid reacts with solid manganese(IV) oxide to form aqueous manganese(II) chloride, liquid water, and chlorine gas.
    d. Liquid pentane (C_5H_{12}) reacts with gaseous oxygen to form carbon dioxide and liquid water.

12. A popular classroom demonstration, solid sodium is added to liquid water and reacts to produce hydrogen gas and aqueous sodium hydroxide. Write the balanced chemical equation for this reaction.

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