AP Chemistry

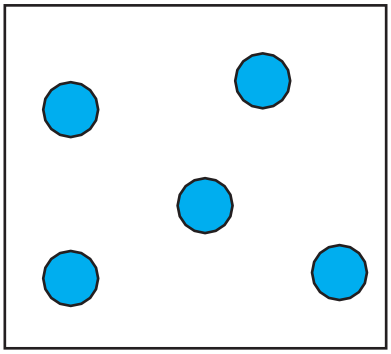
Review Unit

Chemistry is the study of materials and the changes that materials undergo. It is sometimes called the central science because so many naturally occurring phenomena involve chemistry and chemical change.

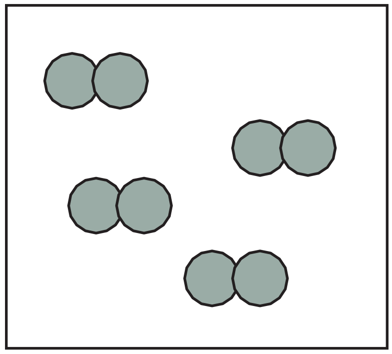
Lesson 1: Chemistry, matter, and ChemiCaL & PhysiCaL Changes

1.1 classification of matter

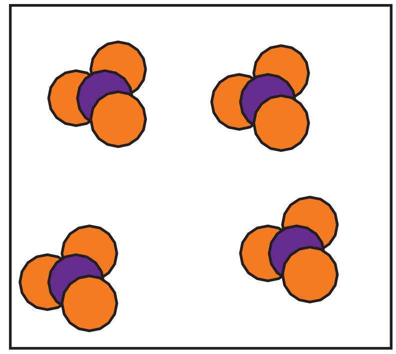
The simplest form of matter is an *element*.
Compounds are combinations of elements that have a definite composition.



atoms of an element



molecules of an element

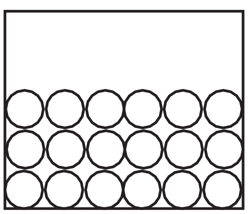


molecules of a compound

Elements and compounds exist as three states of matter: *solids*, *liquids*, and *gases*.

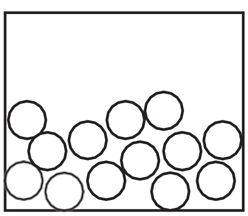
particulate diagrams

Solid



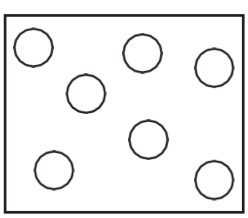
Definite shape
Definite volume
Particles are tightly packed and organized
Particles vibrate gently in fixed positions

Liquid



Indefinite shape (takes shape of container)
Definite volume
Particles are free to slide around one another

Gas



Indefinite shape (takes shape of container)
Indefinite volume (spreads out to fill the space)
Particles are free to move around

All matter has two distinct characteristics: it has mass and occupies space.

1.2 physical and chemical changes and properties

Physical Properties

Physical properties can be observed without changing the identity and composition of the substance.

Examples:
color, odor, density, hardness, solubility, melting point, and boiling point.

Chemical Properties

Chemical properties describe how a substance reacts with other substances

Examples:
acid-base reactions, oxidation and reduction (REDOX), and flammability.

Physical Change

During a physical change, a substance changes its physical appearance, but not its composition.

The most common physical changes are changes of state.

Solid --> Liquid	*Melting*
Liquid --> Gas	*Boiling*
Gas --> Liquid	*Condensing*
Solid --> Gas	*Sublimation*
Gas --> Solid	*Deposition*
Liquid --> Solid	*Freezing*

Chemical Change

In a chemical change, also called a chemical reaction, a substance is transformed into a chemically different substance.

For example, when Hydrogen (H_2) burns in air (O_2), it undergoes a chemical change because it combines with oxygen to form water (H_2O).

$$H_2 + O_2 \dashrightarrow H_2O$$

Chemical changes are often accompanied by observable changes such as color changes and energy changes.

There is a very important distinction to be made between these two types of change that you will encounter in a later unit. For now, note the difference between intermolecular forces and intramolecular forces and how they relate to physical and chemical changes:

During physical changes, the *intermolecular forces*, or IMFs, (the forces between particles) are disrupted. Boiling water separates one water molecule (H_2O) from another water molecule but does not break any individual water molecule apart.

During chemical changes the *intramolecular forces* (the forces within substances) are disrupted. During the electrolysis of water, one water molecule (H_2O) splits up to form O_2 and H_2 atoms. Individual water molecules do break apart.

task 1.2

1. 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) Digesting a candy bar

   (e) Explosion of nitroglycerin

2. Which elements exist as diatomic molecules?

Lesson 2: numbers

2.1 scientific notation

$$C \times 10^{n}$$

+ n represents a large number

- n represents a small number

C is some number: $1 \leq C < 10$

n is an integer
this is the number of times the decimal point moves

Examples

$760000.0 = 7.6 \times 10^{5}$

$0.0035 = 3.5 \times 10^{-3}$

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task 2.1

1. Convert the following numbers to scientific notation.

   (a) 102,500,000

   (b) 0.0000656

   (c) 0.18

   (d) 91,380,000,000

   (e) 1,020

2. Convert the following scientific notation numbers to standard notation.

   (a) 3.70×10^{6}

   (b) 1.92×10^{-2}

   (c) 9.18×10^{-5}

   (d) 1.2×10^{6}

   (e) 7.91×10^{1}

2.2 si units

You are expected to know the following common SI (System International) units and prefixes are given below.

Note: pay close attention to the capitalization of the symbols. M and m are two very different things!

Base quantity	*Unit*	*Symbol*
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s
Amount of substance	Mole	mol
Temperature	Kelvin	K

Prefix	*Symbol*	*Meaning*
Giga	G	10^9
Mega	M	10^6
Kilo	k	10^3
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	μ	10^{-6}
Nano	n	10^{-9}
Pico	p	10^{-12}

Other conversions you may come across in practice

1.00 m = 1.094 yd

1.000 mile = 1760 yd

1.000 kg = 2.205 lbs

1.00 in = 2.54 cm

2.3 derived units

All other units can be derived from base quantities.

Common units for volume: liters (L)
milliliters (mL)

000 mL = 1.000 cm^3

1.000 L = 1000. mL = 1000. cm^3

Density is the ratio of the mass to volume.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Pay attention what units are used when calculating density.

Units for density could be g/L or g/cm^3, etc.

2.4 dimensional analysis

One unit can be converted to another unit by using a conversion factor.

The conversion factor is derived from the equivalence statement of the two units.

For example, in the equivalence of 1.00 m = 100.0 cm, the conversion factor will either be:

$$\frac{100.0\text{ cm}}{1.00\text{ m}} \quad \text{or} \quad \frac{1.00\text{ m}}{100.0\text{ cm}}$$

Remember: The correct choice is the one that allows the cancellation of the unwanted units.

What unit you currently have should appear on the bottom of the conversion factor so that it cancels out.

Example: Convert 289.3 cm to m.

$$289.3\ \cancel{\text{cm}} \quad \times \quad \frac{1.00\text{ m}}{100.0\ \cancel{\text{cm}}} \quad = \quad 2.892\text{ m}$$

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task 2.4

1. Complete the following conversions:

   (a) 0.105 days to seconds

   (b) 0.0550 mi to m

   (c) 0.076 L to mL

   (d) 5.0×10^{-8} m to nm

   (e) 5.850 gal/hr to L/s

   (f) 1.55 kg/m^3 to g/L

2.5 temperature

There are three scales used to measure temperature: Celsius (oC), Fahrenheit (oF) and Kelvin (K).

The following conversion factors will be useful; you must know how to convert between Kelvin and Celsius.

Celsius to Kelvin	*Kelvin to Celsius*
Temperature in oC + 273	Temperature in K - 273
Celsius to Fahrenheit	*Fahrenheit to Celsius*
(1.8 (Temperature in oC)) + 32	(Temperature in oF - 32) / 1.8

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task 2.5

1. Convert the following temperatures from one unit to the other.

   (a) 398 K to oC

   (b) 167 K to oF

   (c) 23 oF to oC

   (d) -13.2 oC to K

   (e) 262 oC to oF

2. When discussing a change in temperature, why will it not matter if the change is recorded in Celsius or Kelvin?

3. How does the Celsius scale compare to that of the Fahrenheit scale? Kelvin?

2.6 uncertainty, significant figures and rounding

All calculations on the AP exam require that the rules for significant figures be obeyed.

Significant figures are the meaningful digits in a measured or calculated quantity.

Determining the number of significant figures present in a number:

all nonzero digits are significant
174 has three sig figs
2719 has four sig figs

all middle zeros are significant
1002 hAs four sig figs
10.2 has three sig figs

leading zeros are not significant
0.00**12** has two sig figs
0.**1987** has four sig fig

trailing zeros are significant if they are to the right of a non-zero digit and there is a decimal in the number

18.0 has three sig figs
0.**8470** has four sig figs

calculations with significant figures

Adding and subtracting
When adding or subtracting, limit the answer to the same number of decimal places that appear in the original data, based on the fewest number of decimal places.

General Strategy	*Example*
Perform the calculation	14.1 – 0.1983 = 13.9017
The 14.1 has the fewest decimal places, the tenths place. Your answer will be rounded to the tenths place.	13.9017 = 13.9

Multiplying and dividing
When multiplying or dividing. Limit the answer to the same number of significant figures that appear in the original data with the fewest number of significant figures.

General Strategy	*Example*
Perform the calculation	189.0 / 1.8 = 157.5
189.0 has four sig figs and 1.8 has two, so your answer will be rounded to two sig figs	157.5 = 160

Combined calculations
Lab measurements were performed to determine the density of an unknown liquid. The following data was obtained in the lab:

Mass of empty graduated cylinder	10.05g
Mass of graduated cylinder + unknown liquid	91.59g
Volume of unknown liquid	88.3 mL

General Strategy	*Example*
Determine the mass of the liquid:	91.58 g – 10.05 g = 81.54 g (two decimal places)
Calculate the density of the unknown liquid once the mass and volume of the liquid are known.	D = m/V D = 81.54 g / 88.3 mL D = 0.9234428086 g/mL
Round to the appropriate number of sig figs. The volume is given as 88.3 mL (three sig figs) and the calculated mass is 81.54 g (four sig figs)	0.923 g/mL

task 2.6

1. Determine the number of significant figures in the following measurements.

(a) 358 kg

(b) 0.054 s

(c) 6.3050 cm

(d) 0.0105 L

(e) 7.0500 x 10^{-3} m^3

(f) 1010. g

2. Use a calculator to carry out the following calculations and record the answer to the correct number of significant figures.

(a) 12.0050 + 9.05

(b) 157.2 – 19.789

(c) 6.32 x 10^3 x 0.1050

(d) 0.0577/0.753

(e) 320.5 – (6104.5/2.3)

(f) (0.045 x 20,000.0) + (2812 x 12)

2.7 accuracy, precision, and error

Accuracy is how close a measurement is to the true value of the quantity that is measured.
Precision is how closely two or more measurements of the same quantity are with one another.

Percent error

Data collected during experiments will often differ from the accepted, published, or actual value.

You express the accuracy of your data using percent error:

$$\text{Percent Error} = \left| \frac{(\text{Actual Value} - \text{Calculated Value})}{\text{Actual Value}} \right| \times 100\%$$

Example

A student must determine the density of a piece of metal during a lab experiment. His calculated value is 2.8 g/mL. The actual density of the metal is 2.6 g/mL. Determine the percent error of the student.

General Strategy	*Example*
Determine the actual value and the calculated value for density according to the problem.	Percent error $= \left\| \frac{(2.6 - 2.8)}{2.6} \right\| \times 100\%$
Determine the answer	Percent error = 7.69230 %
Round your answer to the correct number of sig figs	2.6 - 2.8 = - 0.2 (1 sf) 0.2 / 2.6 = answer in 1 sig fig 8% error

task 2.7

1. Two students determine the percentage of lead in a sample as a laboratory exercise. The true percentage is 22.52%. The students' results for the three trials are as follows:

   Trial 1: 22.52, 22.48, 22.54

   Trial 2: 22.64, 22.58, 22.62

   a. Calculate the average percentage for each set of data, and tell which set is the more accurate based on the average.

   b. Which set is more precise? Explain.

   c. Using your averages for part a, calculate the percent error for each student.

Lesson 3: atomic theory

3.1 key scientists and their contributions to the atomic theory

Antoine Lavoisier	***Law of Conservation of Mass*** Mass is neither created nor destroyed in a chemical reaction.
Joseph Proust	***Law of Definite Proportion*** The same compound always contains exactly the same proportion of elements by mass.
Dmitri Mendeleev	***Periodic Table of Elements*** He was the first to conceive the modern Periodic Table of Elements. Insisted certain spots of the table be left blank until the actual element is found that matched the predicted properties. This was done to preserve the elements with similar properties called groups or families.
John Dalton	***Law of Multiple Proportion*** When two elements form a series of compounds, the ratios of the masses of the second element that combine with the first element can always be reduced to small whole numbers. ***Dalton's Atomic Theory*** 1. All elements are made up of tiny particles called atoms. 2. The atoms of a particular element are identical. Different elements have different kind of atoms. 3. Atoms cannot be created or destroyed. 4. Chemical compounds are formed when different kinds of atoms combine together. A particular compound always has the same relative numbers and types of atoms. 5. Chemical reactions deal with the rearrangement of the atom, which changes the way they are combined together. There is no change to the atoms themselves in a chemical reaction. ***Dalton's Law of Partial Pressures*** The total pressure exerted by a gaseous mixture is equal to the sum of partial pressures of each individual component in a gas mixture. (Ptotal = PA + PB + PC + …)

J.J. Thomson	He measured the charge to mass ratio of an electron using a ***Cathode Ray Tube*** ***Plum Pudding Model*** Electrons are embedded in a cloud of protons.
Robert Milliken	***Milliken Oil Drop Experiment*** Found the mass (9.11×10^{-31} kg) and the charge (1.6×10^{-19} C) of an electron when balancing the electric force and gravitational force of an electron across a set of charged plates.
James Chadwick	Discovered neutrons
Ernest Rutherford	Nuclear Model He performed the famous ***Gold Foil Experiment*** and proposed that the protons and neutrons are in the centre of an atom (nucleus) where the electrons fly around the nucleus. The nucleus is very small and the atom is mainly made of empty space
Neil Bohr	***Bohr Atomic Model*** Electrons are in specific orbits (energy level) around the nucleus, and therefore electrons are quantized. He helped developed the ***Quantum Mechanics Model*** (Electron Cloud Model), that is based on mathematical probabilities. Formulated the ***Aufbau Principle***, which states the electrons fill orbitals starting at the lowest available energy states before filling higher states.
Max Planck	He proposed the light could be viewed as particles as well as waves.

3.2 structure of the atom and the periodic table

Protons, Neutrons, and Electrons

Name	*Charge*	*Mass in amu*	*Position in atom*
Proton	+1	1	nucleus
Neutron	0	1	nucleus
Electron	-1	1/1836	outside of nucleus

Mass Number — A
Atomic Number — Z

$$^{A}_{Z}X$$

Z = atomic number = # of protons (also # of electrons in a neutral atom) A
A = mass number = # of protons + # of neutrons in one atom of an element

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Ions

Ion	*Charge*	*Formed*
Cation:	+	loses 1^{+} electrons
Anion:	-	gains 1^{+} electrons

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task 3.2

1. Determine the number of protons, electrons and neutrons in:

$^{36}_{17}C$

$^{124}_{53}I$

Lesson 4 nomenclature

4.1 symbols

It is important when writing the two letter symbols to ensure that you use a lower case letter for the second letter. This may sound trivial but is very important, for example, Co (cobalt), a metal element, is not the same as CO (carbon monoxide), a gaseous compound made from carbon (C) and oxygen (O).

$$Co \neq CO$$

Monatomic Ions

Cations formed from *metal* atoms have the same name as the metal

Na^{+} sodium ion
Zn^{2+} zinc ion
Al^{3+} Aluminum ion

Cations formed from *transition metals* that form different cations will have the charge indicated by a roman numeral after the name of the metal

Fe^{2+} iron (II) ion
Fe^{3+} iron (III) ion
Cu^{+} copper (I) ion
Cu^{2+} copper (II) ion

Monatomic ***anions*** are names by replacing the ending of the name of the element with –ide

H^{-} hydride ion
O^{2-} oxide ion
N^{3-} nitride ion

4.2 binary ionic compounds

Naming:
Metal + Nonmetal
Cation + Anion –ide ending

Examples:
NaCl = Sodium chloride
$CaBr_2$ = Calcium bromide
Li_2O = Lithium oxide

Formulas:
Balance the charges
Ca^{2+} + F^{-1}
CaF_2

Examples:
Strontium iodide = SrI_2
Potassium chloride = KCl
Magnesium oxide = MgO

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task 4.2

1. Name these binary compounds.

(a) NaCl

(b) SrO

(c) AlN

(d) $BaCl_2$

(e) K_2O

(f) CuO

(g) Cu_2O

2. Convert these names to chemical formulas.

(a) Magnesium nitride

(b) Barium bromide

(c) Aluminum phosphide

(d) Potassium iodide

(e) Lithium chloride

(f) Sodium fluoride

(g) Tin (IV) bromide

4.3 binary covalent compounds

Naming:

1. The name of the element farther to the left of the periodic table is usually written first
   **Exception: oxygen is always written last*

2. If both elements are in the same group in the periodic table, the one having the higher atomic number is named first

3. The name of the second element is given the –ide ending

4. Greek prefixes are used to give the number of atoms of each element

Mono	1
Di	2
Tri	3
Tetra	4
Penta	5
Hexa	6
Hepta	7
Octa	8
Nona	9
Deca	10

Examples:

Cl_2O dichlorine monoxide
N_2O_4 dinitrogen tetroxide
NF_3 nitrogen trifluoride
P_4S_{10} tetraphosphorus decasulfide

task 4.3

1. Write formula or names for the following molecular compounds.

   (a) Dinitrogen tetroxide

   (b) Phosphorous pentachloride

   (c) Iodine trifluoride

   (d) Nitrogen dioxide

   (e) Dihydrogen monoxide

2. Convert the following formulae to names.

   (a) N_2O_5

   (b) PCl_3

   (c) SF_6

   (d) H_2O

   (e) Cl_2O

4.4 binary acids

We will discuss acid in depth in a later unit, but for now be able to recognize an acid as a compound that produces hydrogen ions (H^+) when dissolved in water.

The formula of an acid typically starts with an "H"

Binary Acids are acids where hydrogen combines with a monatomic anion

Hydro + (element stem)ic acid

HCl = Hydrochloric acid
HF = Hydrofluoric acid

4.5 polyatomic ions

Some common polyatomic ions, their charges and formulas are listed below. You will need a

more complete list for the purpose of this class.

Name	***Charge***	***Formula***
Ammonium	1+	NH_{4+}
Carbonate	2-	CO_{32-}
Chromate (VI)	2-	CrO_{42-}
Dichromate (VI)	2-	Cr_2O_{72-}
Hydrogen carbonate	1-	HCO_{3-}
Hydrogen sulfate	1-	HSO_{4-}
Hydroxide	1-	OH^-
Manganate (VII) (permanganate)	1-	MnO_4^-
Nitrate	1-	NO_{3-}
Nitrite	1-	NO_{2-}
Phosphate	3-	PO_{43-}
Sulfate	2-	SO_{42-}
Sulfite	2-	SO_{32-}

Don't forget about the -ate trick!

BO_3^{-3} Borate	CO_3^{-2} Carbonate	NO_3^{-1} Nitrate		
		PO_4^{-3} Phosphate	SO_4^{-2} Sulfate	ClO_3^{-1} Chlorate
		AsO_4^{-3} Arsenate	SeO_4^{-2} Selenate	BrO_3^{-1} Bromate
				IO_3^{-1} Iodate

Group of 3 = XO_3
Borate
Carbonate
Nitrate

Group of 4 = XO_4
Phosphate
Arsenate
Sulfate
Selenate

Group of 3 = XO_3
Chlorate
Bromate
Iodate

In oxyanions, their names will follow the trend below:

Hypo – ite -2 oxygens
-Ite -1 oxygen
-Ate base oxygen
Per – ate +1 oxygen

Some oxoanions contain hydrogen and are named accordingly:
Example, HPO_4^{2-}, hydrogen phosphate.

task 4.4 -4.5

1. What are the formulae for the following ionic compounds?

   (a) Ammonium nitrate

   (b) Copper (II) bromide

   (c) Copper (I) bromide

   (d) Zinc hydrogen sulfate

   (e) Aluminum sulfate

   (f) Sodium perchlorate

   (g) Copper (II) iodite

2. Convert the following formulae to names.

   (a) $NaNO_3$

   (b) $KMnO_4$

   (c) $CaCO_3$

   (d) $CuSO_4$

   (e) $Cu2SO_4$

   (f) KNO_2

   (g) $LiClO_4$