MATTER

✱ Definition: Anything which has mass and takes up space.

♦ Question: Can it be separated by physical means?

✔YES → MIXTURE ✘NO → PURE SUBSTANCE

MIXTURE

✱ Definition: Matter which contains ***2 or more elements and/or compounds mixed together***.

♦ Question: Is it uniform throughout?

✔YES → HOMOGENEOUS MIXTURE ✘NO → HETEROGENEOUS MIXTURE

HOMOGENEOUS MIXTURE

(also called a SOLUTION)

✱ Definition: A mixture which ***has a uniform*** composition throughout the entire sample.

☞ Composition ***does not vary*** throughout the mixture.

☞ Does not exhibit ***Tyndall Effect***

EXAMPLES

HETEROGENEOUS MIXTURE

✱ Definition: A mixture which ***does not have a uniform*** composition throughout the mixture.

☞ Composition ***does vary*** throughout the mixture.

☞ Two Types: ***suspension*** and ***colloid***

☞ Exhibits ***Tyndall Effect***

EXAMPLES

PURE SUBSTANCE

✱ Definition: Matter which contains a ***single element or compound*** throughout the entire sample.

♦ Question: Can it be separated by chemical means?

✔YES → COMPOUND ✘NO → ELEMENT

COMPOUND

✱ Definition: Made up of ***2 or more different types of atoms*** bonded together the same way throughout the entire sample.

☞ Smallest unit is called a ***molecule*** or ***formula unit***.

☞ ***molecules*** and ***formula units*** are made up of ***atoms***.

EXAMPLES

ELEMENT

✱ Definition: Made up entirely of ***1 type of atom***

☞ Smallest unit is called an ***atom***.

EXAMPLES

Atomic Structure Notes

Subatomic Particle	Abbreviation	Charge	Location	Mass (a.m.u.)
Proton	p^+	+1	Nucleus	1
Neutron	n^0	0	Nucleus	1
Electron	e^-	-1	Orbitals Surrounding the Nucleus	Negligible, Almost 0

Nucleus- extremely dense center of the atom, consists of protons and neutrons, contains almost all of the mass and virtually none of the volume of the atom

Atomic Symbol- the one or two letter abbreviation of an element

Atomic Number- the identifying number of a specific atom, equal to the number of protons for that element

Mass Number- the number of protons and neutrons

Nuclear Symbol- Mass #

Atomic Symbol

Atomic #

Isotope- atoms of the same element with a different number of neutrons

Ion- an electrically charged atom

Cation- a positively charged ion resulting from losing an electron

Anion- a negatively charged ion resulting from gaining an electron

Rule for Electron Configuration

Aufbau- States that electrons fill beginning with the lowest energy levels first.

Hund's Rule- States that electrons fill orbitals of the same energy (degenerate orbitals) by adding one electron to each orbital all with the same spin and then doubling up.

Pauli Exclusion Principle- States that any orbital can only hold a maximum of two electrons with opposite spins.

Chapter 4 Notes

Periodic Table of Elements- chart of all the chemical elements arranged with columns (groups) of elements with similar chemical properties similar

Groups-columns of the periodic table
Periods- rows of the periodic table
Alkali Metals- Group 1 elements
Alkaline Earth Metals- Group 2 elements
Transition Metals- Group 3-12 elements
Halogens- Group 17 elements
Noble Gases- Goup18 elements
Main Group Elements- Group 1, 2, 13-18 elements

Valence Electrons- an electron in the outer most energy levels of an atom, where it can participate in bonding

Octet Rule- the tendency for main group elements to gain or lose electrons so that their s and p orbitals are full with 8 electrons

	Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group18
Electron Ending	s^1	s^2	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6
Valence Electrons	1	2	3	4	5	6	7	8
Gain or Lose electron	Lose 1 electron	Lose 2 electrons	Lose 3 electrons	Lose or Gain 4 electrons	Gain 3 electrons	Gain 2 electrons	Gain 1 electron	Gain or Lose 0 electrons
Common Charge	+1	+2	+3	+/- 4	-3	-2	-1	0

Ionic Compounds (salts)- compounds that form due to the transfer of electrons, usually bonding a metal to a non-metal, smallest unit is called a formula unit.

Covalent (Molecular) Compounds- compounds that form due to the sharing of electrons between elements, usually bonding non-metals to non-metals, smallest unit is called a molecule.

Main Group Elements (Group 1,2, 13-18) readily form ions to have an electron configuration like the stable Noble Gases. Main Group Elements become isoelectronic with the nearest Noble Gas.

Trends of the Periodic Table

Atomic Radius- the radius of a neutral atom, measured by 1/2 the distance between nuclei when bonded to another like atom.

Trend

Increases down a group because an extra principle energy level is added

Decreases across a period because an extra proton and electron is added for each atom, which increases the attraction of the nucleus to the surrounding electrons

Ionization Energy-the energy it takes to remove one electron from a neutral atom in gaseous state

High ionization energy- very difficult to remove an electron

Low ionization energy- very easy to remove an electron

Trend

Decreases down a group, increases across a period

Note:

Smaller atoms hold their outer electron more tightly, making it harder to remove an electron, therefore it takes more energy to remove them

Larger atoms do not hold their outer electrons as tightly, making it easier to remove an electron, therefore it takes less energy to remove them

Electron shielding- the reduction of the attractive forces between the positively charged nucleus of an atom and the outermost electrons

Electronegativity- the tendency of an atom to attract bonding electrons to itself when bonded to another atom.

Note:

The smaller the atom, the more it will pull bonded electrons towards it nucleus, therefore higher electronegativity.

Noble Gases are not included in this trend because they are unreactive

Trend	Down a Group	Across a Period	Include Noble Gases in trend	Highest Value
Atomic Radius	Increases	Decreases	Yes	Francium
Ionization Energy	Decreases	Increases	Yes	Helium
Electronegativity	Decreases	Increases	No	Fluorine

Trend	Atomic Radius	Ionization Energy	Electronegativity
Smaller Atoms	↓ Lower	↑ Higher	↑ Higher
Larger Atoms	↑ Higher	↓ Lower	↓ Lower

Ionic Radius- the radius of an atom after it has gained or lost an electron

Anion- a negative ion, an atom that has gained electrons
Cation- a positive ion, an atom that has lost electrons

The radius of a cation will be much smaller than the original neutral atom
The radius of an anion will be much bigger than the original neutral atom.

Isoelectronic- an atom or ion with the same number of electrons as another ion or atom

For isoelectronic ions...

the greater the positive charge the smaller the radius
the greater the negative charge the larger the radius

Reactivity

Metals- the bigger the atom, the more reactive

Large atomic radius
Smaller ionization energy-easier to take electrons, because they are excellent electron givers

Non metals- the smaller the atom, the more reactive

Excellent electron takers
Smaller atomic radii
High ionization energy- very hard to take electrons

Noble Gases are not included in this trend because they are unreactive

Group 1- Alkali Metals- soft metals, highly reactive (+1)

Group 2- Alkaline Earth metals, hard metals, quite reactive (+2)
Group 3-12 Transition Metals- d-block metals, somewhat reactive (variable positive charges)
Inner Transition metals- Lanthinides/Actinides, f block metals
Halogens Group 17- distinct colors, 2 gases, 1 liquid, 1 sold, poisonous, quite reactive
Noble Gas- unreactive (inert) colorless gases, unreactive because of stable electron configurations

Balancing Equations Notes

Background Info
Diatomic Elements- Elements that exist in pairs when not bonded to any other different element.
BrINClHOF

Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

Acids

Hydrochloric Acid- HCl	Sulfuric Acid- H_2SO_4
Nitric Acid- HNO_3	Phosphoric Acid H_3PO_4
Acetic Acid- $HC_2H_3O_2$	Carbonic Acid- H_2CO_3

Balanced Chemical Equations
1. Word Equation

Zinc added to hydrochloric acid yields zinc chloride and hydrogen gas
Zinc + hydrochloric acid → zinc chloride + hydrogen gas
Reactants Products

2. Unbalanced Equation
-Correctly write chemical formulas:

$Zn + HCl \rightarrow Zn\,Cl_2 + H_2$

1atom Zn	1 atom Zn
1 atom H	2 atom H
1 atom Cl	2 atom Cl

- Never change subscripts of correctly written formulas.
-Violates law of Conservation of Matter

3. Balanced Chemical Equation

- Both sides of the equation must have the same number of atoms of each element.
- To balance the equation, add coefficients in front of correctly written formulas.

$Zn + HCl \rightarrow Zn\,Cl_2 + H_2$

1 atom Zn	1 atom Zn
1 atom H	2 atom H
1 atom Cl	2 atom Cl

$Zn + 2HCl \rightarrow Zn\,Cl_2 + H_2$

1atom Zn	1 atom Zn
2 atom H	2 atom H
2 atom Cl	2 atom Cl

Writing word equations procedure

1. Identify the ions. (Use ion sheets)
2. Write the correct chemical formulas for all reactants and products.
3. Count atoms of each type on each side. If a polyatomic ion appears on both sides, treat it as a single unit.
4. Start with the elements that appear only once and then move on to others. Other elements that appear several times will balance themselves out.
5. Place coefficients in front of compounds and/or free elements as necessary.

Type of Reaction	Abbrev.	Reactants	Products	Example
Synthesis	SYN	2 Free Elements	1 Compound	$A + X \rightarrow AX$
Decomposition	DEC	1 Compound	2 Free Elements	$AX \rightarrow A + X$
Single Replacement	SR	1 Element and 1 Compound	1 Element and 1 Compound	$A + BX \rightarrow AX + B$ or $AX + Y \rightarrow AY + X$
Double Replacement	DR	2 Compound	2 Compound	$AX + BY \rightarrow AY + BX$
Combustion	COMB	Hydrocarbon ($C_xH_YO_z$) and Oxygen Gas (O_2)	Water (H_2O) and Carbon Dioxide (CO_2)	$C_xH_YO_z + O_2 \rightarrow H_2O + CO_2$

Key

A and B = positive ions X and Y = negative ions

Combustion reactions

$C_xH_YO_z + O_2 \rightarrow CO_2 + H_2O$

1. Balance carbon first.
2. Balance hydrogen next.
3. Then balance oxygen.
4. If there is an odd number of oxygen on one side multiply all coefficients by two

Predicting products procedure

1. Based on reactants determine reaction type.
2. Identify the ions.
3. Determine which elements/ions combine with each other or separate from each other.
4. Write the correct formulas for products (use ion sheets).
5. Balance equation with coefficients only

Reminders

1. Double check chemical formulas!!!
2. Like charges will be exchanged in replacement reaction.
3. Free elements will be monatomic with zero charge except Br, I, N, Cl, H, O, and F which will exist by themselves as Br_2, I_2, N_2, Cl_2, H_2, O_2, and F_2.
4. Treat H_2O as HOH (an ionic combination of the H^+ and OH^- ions), in single replacement (SR) and double replacement (DR) reactions
5. Treat polyatomic ions as single ion. Do not break up into constituent elements.
6. When balancing an unbalanced equation, only change coefficients. Do not mess with chemical formulas.
7. Never use "In Between" coefficients.

Chapter 6 notes

	How Bonded…	Made up of…	Representative Particle name	Abbrev.
Elements	N.A. Not bonded.	1 element only	Atom	at.
Ionic Compounds	Ionic Bonds-transfer/ exchange of electrons	Usually a metal and non-metal	formula unit	f.u.
Covalent/ Molecular Compounds	Covalent bonds- sharing of electrons	Usually non-metals	Molecule	m.c.

The Mole

Avogadro's Number = $6.02 * 10^{23}$ 1 mole = 6.02×10^{23} r.p.

Examples:

1 mol Cu = 6.02×10^{23} atoms of Cu

1 mol H_2O = 6.02×10^{23} molecules of H_2O

1 mol NaCl = 6.02×10^{23} formula units of NaCl

Molar Mass

Molar Mass: The mass in grams of 1 mole (6.02×10^{23} particles) of any substance.

-For Any Element- The atomic mass of that element expressed in grams.

-For Any Compound- The sum of all the atomic masses of a compound expressed in grams.

-Conversion factor of # of grams per 1 mole

	Atomic Mass (1 particle)	Molar Mass (6.02×10^{23} particles)
Fe	55.8 amu	55.8 grams
$Fe_3(PO_4)_2$	291.6 amu	291.6 grams

Fe: 3 x 55.8 g = 111.6 g 111.6 g + 62.0 g + 128.0 g = 291.6 g

P: 2 x 31.0 g = 62.0 g

O: 8 x 16.0 g = 128.0 g $\boxed{291.6 \text{ g } Fe_3(PO_4)_2 = 1 \text{mole } Fe_3(PO_4)_2}$

Moles → r.p. (atoms, molecules, formula units) # moles x $\frac{6.02 * 10^{23} \text{ r.p.}}{1 \text{ mole}}$ = r.p.	r.p. → moles # r.p. x $\frac{1 \text{mole}}{6.02 * 10^{23} \text{ r.p.}}$ = moles
Moles → Grams # moles x $\frac{\text{molar mass (g)}}{1 \text{ mole}}$ = mass (g)	Grams→ Mole #g x $\frac{1 \text{ mole}}{\text{molar mass (g)}}$ = moles

Grams ← Molar Mass → Moles ← Avogadro's #→ Particles

Percent Composition notes

Empirical Formula- The most reducible form of a molecular formula

Hydrate- an ionic compound with a precise number of water molecules in its crystal structure (not bonded)

Anhydrous- an ionic salt that has had its water molecules removed

Formula → % Composition

1. Divide contribution by molar mass (multiply by 100)

$$\% \text{ composition of element} = \frac{\text{contribution of element (in grams)}}{\text{Molar Mass (in grams)}} \times 100 =$$

Example: $C_6H_{12}O_6$

6 x12.0 g C = 72.0 g C	72.0 g C/ 180.0 g C x 100 = 40.0 % C
12 x 1.0 g H = 12.0 g H	12.0 g H/ 180.0 g H x 100 = 6.67 % H
6 x 16.0 g O = <u>96.0 g O</u>	96.0 g O/ 180.0 g O x 100 = 53.33 % O
180.0 g $C_6H_{12}O_6$	