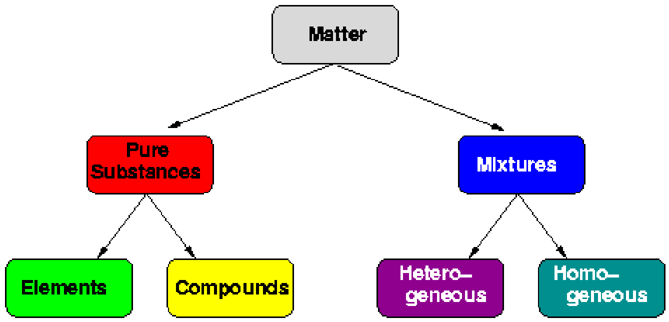
UNIT I: **CHEMICAL REACTIONS** ***Text: Chapters 5-8*** SCIENCE 1206

Chemistry: the study of matter, its properties and its changes.
Matter: anything that has mass and takes up space (energy is **not** matter).
The three states of matter: **solid, liquid and gas.**

Classification of Matter as Pure Substances or Mixtures:



Pure Substances: have constant composition; all the particles that make up the substance are the same.

1. **Elements: -** the simplest form of matter that can exist under normal conditions.
 - composed of only *one kind of atom.* (eg. C, Mg)
 - cannot be broken into simpler substances by chemical means (heat/electricity).
 - combine to form other substances.

2. **Compounds:** -substances composed of *two or more different kinds of atoms* eg. H_2O, NaCl.
 - can be broken down into simpler substances by chemical means.

Mixtures: have variable composition: composed of 2 or more pure substances.

1. **Homogeneous Mixtures:** **solutions –** have only one visible component
eg. tap water, air, sugar solution (sugar + water)

2. **Heterogeneous Mixtures:** **mechanical mixtures** – have 2 or more visible components.
eg. sand in water, vegetable soup

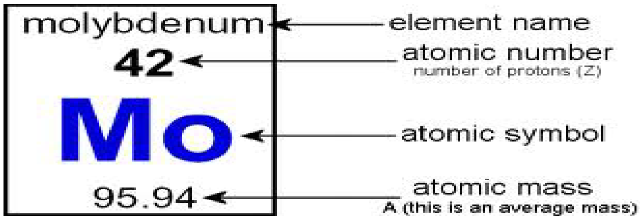
Properties and Changes of Matter:

1. **Physical Property:** Characteristics of matter, used to identify substances.
eg. state at room temperature, boiling and melting points, color, solubility, mass, electrical conductivity, odour and luster.
2. **Physical Change:** A change in the size or form of a substance that does not change its composition. **eg.** cutting, bending, changes in state: boiling, melting, condensing, and solidifying.
3. **Chemical Property:** Characteristic of matter that can be observed when matter undergoes a change in composition (chemical reaction): describes "how it reacts"
eg. butane reacts with oxygen to produce carbon dioxide and water.
4. **Chemical Change:** A chemical reaction; a change in which at least one or more new substances (**products**) are formed. The products have different properties from the starting substances (**reactants).**
eg. $Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$ The rust produced has completely different properties from iron and oxygen.

Evidence of Chemical Change:
change in **color, odor, energy** (temperature change, light)
change in **state:** **bubbles = new gas produced**
precipitate = new solid produced

ELEMENTS & THE PERIODIC TABLE

The information provided by the periodic table for each element.
Ex1



All elements are classified as metals or nonmetals

All elements are classified as metals or nonmetals, depending on their properties.

PROPERTY	METALS	NONMETALS
LUSTRE	shiny	dull
MALLEABILITY	malleable (bendable)	brittle
CONDUCTIVITY OF HEAT & ELECTRICITY	good conductors	poor or nonconductors
STATE AT ROOM TEMPERATURE	all solids except mercury, Hg = liquid	most are gases, some are solids and bromine, Br = liquid
REACTIVITY WITH ACID	mostly yes	no
LOCATION (PERIODIC TABLE)	**left** of staircase line	**right** of staircase line

METALLOIDS (Semimetals)
- elements that have some properties of metals and some properties of nonmetals.
- include all elements on either side of the staircase line **except Al and At.**
- also includes one form of Carbon, **graphite**, which is dull and brittle (nonmetal), but is a good conductor of electricity (metal).

CHEMICAL FAMILIES (GROUPS): Groups of elements in the same vertical column that have similar physical and chemical properties.

1. **Alkali Metals: -** Group 1, IA
   - show metallic properties (see table above)
   - highly reactive, especially with water; reactivity increases going down the group
   - Cs & Fr are the most reactive metals
   - form compounds that are mostly white solids and are very soluble in water.

2. **Alkaline Earth Metals: -** Group 2, IIA
   - show metallic properties (see table above)
   - less reactive than alkali metals; reactivity increases going down the group.
   - form compounds that are often insoluble in water.

Note: Metals from both group 1 and group 2 react with water to form alkaline (basic) solutions.

3. **Halogens:** - Group 17, VIIA

- show nonmetallic properties (see table above)
- reactivity decreases going down the group: F is the most reactive nonmetal
- react with most metals to produce salts (ionic compounds)
- react with hydrogen to form compounds that dissolve in water to form acids

4. **Noble Gases: -** Group 18, VIIIA
   - show nonmetallic properties
   - extremely low chemical reactivity
   - generally do not form compounds (in 1962 one was synthesized in BC – xenon hexafluoroplatinate)

SERIES OF ELEMENTS:

1. **Representative Elements:** A groups (1A to 8A) or groups 1, 2, 13 –18
2. **Transition Elements:** B groups or groups 3 – 12

HYDROGEN:- the lightest element and most abundant element in the universe
- doesn't really belong to any group
- it sometimes behaves like an alkali metal, sometimes like a halogen and at other times in its own unique way ie. as an acid

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PERIODS: horizontal rows of the periodic table

THE ATOM:

- The basic building block of all matter
- The smallest particle of an element that retains the properties of that element
- Electrically neutral: # of positive charges = # of negative charges

- Composed of 3 types of subatomic particles:

Diagram:

PARTICLE	SYMBOL	RELATIVE CHARGE	ACTUAL MASS (g)	LOCATION
Proton	p^+	1^+	1.67×10^{-24}	nucleus
Neutron	n^o	0	1.67×10^{-24}	nucleus
Electron	e^-	1^-	9.11×10^{-28}	orbital

Atomic Number:- identifies the element
- equal to the number of protons in the nucleus
- since atoms are electrically neutral, # of protons = # of electrons

Mass Number: - # of protons + # of neutrons
- protons and neutrons account for most of the mass of the atom

How many protons, electrons, and neutrons are in the following atoms?

Element	Atomic #	Mass #	#p	#e	#n
Carbon (C)		12			
Carbon (C)		13			
Magnesium (Mg)		24			
Magnesium (Mg)		26			
Sodium (Na)		23			
Chlorine (Cl)		36			

Complete the following chart

Atomic Number	Mass Number	Number of protons	Number of electrons	number of neutrons	Element Symbol
8				9	
		11		10	
	40		20		
	88	38			

Quantum Mechanics Theory of the Atom:

According to this theory, an electron with a specific energy occupies a region in space (**orbital**) or electron energy level.
Diagram:

Electron Energy Diagrams of Atoms:

- An energy level represents a specific value of energy of an electron and corresponds to a general location
- The number of occupied energy levels in any atom is normally the same as the **period number** in which the atom appears
- for the first 3 energy levels, the maximum number of electrons that can be present are 2, 8 and 8 in order of increasing energy (increasing distance from nucleus)
- a lower energy level is filled with electrons to its maximum before the next level is started.
- the electrons in the highest (outermost) occupied energy level = **valence electrons,** which is the same as the **group number** (for group A elements)

Energy level diagrams for atoms:
Draw energy level diagrams for the following atoms: magnesium and chlorine.

STABLE ATOMS

- have low chemical reactivity
- include noble gases, all of which have 8 valence electrons (except He, which has 2)

- other atoms can become more stable by reacting and changing the number of their electrons, thereby attaining the same stable electron configuration (structure) of the nearest noble gas:
    - atoms can follow one of two rules:

  a) Octet Rule:- atoms attempt to obtain 8 valence electrons
  - includes most atoms

  b) Duet Rule: - atoms attempt to obtain 2 valence electrons - includes H, Li and Be

- one way atoms can achieve a stable octet or duet is by forming *ions*

IONS

- an atom or group of atoms that gain electrons attain an overall negative charge as a ion.
- an atom or group of atoms that lose electrons attain an overall positive charge as a ion.
    - single atoms: form simple ions (*monatomic ions*)
    - group of atoms: form complex ions (*polyatomic ions)*

Energy level diagrams for ions
Examples: Draw energy level diagrams for the following atoms: magnesium ion and chloride ion.

Example: Sodium metal and chlorine gas react to produce NaCl, a very stable and unreactive substance, compared to Na (alkali metal) or Cl (halogen).
They do so by first forming ions

Summary: When sodium metal and chlorine gas react, the sodium atoms each lose one electron to a chlorine atom. In so doing the atoms form ions of opposite charge:

Cations: - Ions which form when atoms *lose* one or more valence electrons
- *metal* atoms form cations
- for group A atoms, # of valence electrons lost = group # = charge on ion
- have a *positive charge* because they have more protons than electrons

Anions: - Ions which form when atoms *gain* one or more valence electrons
- *nonmetal* atoms form anions
- for group A atoms, # of valence electrons gained = 8 – group # = charge on ion
- have a *negative charge* because they have more electrons than protons

Note:
Remember that families of elements have similar chemical and physical properties. These families of elements will gain, or lose, specific numbers of electrons to attain a stable 'noble gas like' electron arrangement. All elements in group IA, for example will lose one electron to be like the nearest noble gas. The other families are as follows.

Group	**Gain or Loss of electrons to become an Ion**
Group IA (alkali metals)	- lose of one electron to become an ion
Group IIA (alkaline earth)	- lose two electrons to become an ion
Group IIIA	- lose three electrons to become an ion
Group VA	- gain three electrons to become an ion
Group VIA	- gain two electrons to become an ion
Group VIIA (halogens)	- gain one electron to become an ion
Group VIIIA (noble gases)	- Do not gain or lose electrons to become an ion

Valence Electrons- are the number of electrons in the last energy level of an element.

Group A	1A	2A	3A	4A	5A	6A	7A	8A
# valence electrons	1e-	2e-	3e-	4e-	5e-	6e-	7e-	8e-

Naming Ions:

Cations: element name + the word "ion" eg: Na^+ = sodium ion

Anions: stem of element name + ide + the word "ion"

eg: Cl, chlorine becomes Cl^- = chloride ion

P, phosphorus becomes P^{3-} = phosphide ion

O, oxygen becomes O^{2-} = oxide ion

Note:

1. Both cations and anions are more stable than the atoms from which they form since these ions attain the same stable electron configuration as the nearest noble gas.
2. Boron, carbon and silicon do not tend to form ions (they instead share electrons with other atoms)
3. The noble gases do not form ions since they are already stable (have filled orbitals)
4. Hydrogen can form a cation or an anion:
   - Cation: H^+ , hydrogen ion has 1 proton but no electrons
   - Anion: H^- , hydride ion has 1 proton and 2 electrons

Ions

Name of Ion	Symbol	Atomic number	# of protons	# of electrons	# of neutrons
chloride					
		11			
			15		
Zinc Ion					
		34			
			12		

IONIC COMPOUNDS

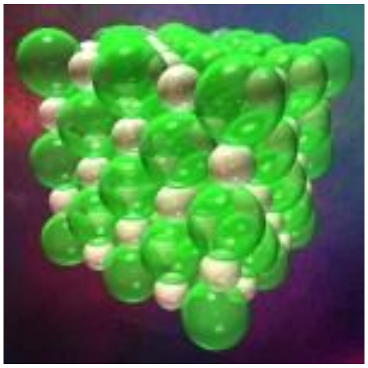
- All are solids at SATP (Standard Ambient Temperature and Pressure) of 25^{o}C and 100 kPa.

- When they dissolve in water, they form *aqueous* solutions that:
  - are colored or colorless
  - conduct electricity ie. they are **electrolytes**

- These compounds form after an electron transfer:
  - usually *from a metal to a nonmetal*
  - the resulting ions (cations and anions) are attracted to each other (since they are oppositely charged) and they form **ionic bonds**

  http://www.youtube.com/watch?v=xTx_DWboEVs&feature=related

- Together all of the ions present form an **ionic crystal lattice** in which the net charge is zero



Eg: (1) in a sample of sodium chloride, NaCl,
for every Na^+ ion there is one Cl^- ion

Eg: (2) in a sample of calcium chloride, $CaCl_2$, for every Ca^{2+} ion there are 2 Cl^- ions

Formula Unit: an expression of the simplest whole number ratio of cations to anions

Eg: NaCl 1:1 Ratio of Na^+ : Cl^-

$CaCl_2$ 1:2 Ratio of Ca^{2+} : $2Cl^-$

http://www.youtube.com/watch?v=QqjcCvzWwww&feature=related

Types of Ions and Ionic Compounds

A. Monatomic Ions (Simple Ions)

- Single atoms that have lost or gained one or more electrons
- Form *Binary Ionic Compounds* (2 simple ions)

Binary Ionic Compounds
Metals and nonmetals combine to form ionic compounds by transferring electrons. The result is a compound that is electrically neutral. The sum of the charges on the positive ion equals the sum of the charges on the negative ions. Ex Na^+ Cl^-
Example 1: Name the compound and chemical formula for the combination of aluminum (metal) and chlorine (nonmetal).

Solution: Aluminum chloride

Note that naming is according to IUPAC Standards (International Union of Pure and Applied Chemistry). The metal name stays the same while the nonmetal assumes an **ide** ending.

To write the chemical formula, first write the symbol and ionic charge for each element.
Al -- 3+
Cl -- 1-
Now ask yourself how many ions of each type would you need so that the net charge is 0. 3 Cl^- will balance 1 Al^{3+}. Thus the formula is $\mathbf{AlCl_3}$

Ex2: Write the chemical formulas.

a) Silver chloride

c) Aluminum oxide

b) Barium fluoride

d) Magnesium nitride

Extra Examples:

Ex 3: Write the names for the following ionic compounds.

a) K_3N

c) Li_2O

b) $BaCl_2$

d) Sr_3P_2

Extra Examples:

B. Polyatomic Ions (Complex Ions)

- Cations or anions composed of a group of atoms with a net positive or negative charge
- When compounds containing these ions are dissolved in water, the polyatomic ion is a tightly bonded groups of atoms that behave as a unit

Eg. NH_4^+	NO_2^-	NO_3^-	CO_3^{2-}
Ammonium ion	Nitrite ion	Nitrate ion	Carbonate ion

See page 196 for a table of some common polyatomic ions.

Note that rules for naming are the same as those for binary ionic compounds. Just remember that the complex ion remains as one.

Ex1: Write the name and chemical formula for the compound formed by sodium and a carbonate ion.

Solution:

1+	2-
Na	CO_3

This gives the formula Na_2CO_3 which is sodium carbonate**.**

Ex2: Write the chemical formulas.

a) Ammonium nitrate

c) potassium carbonate

b) Barium phosphate

d) Sodium sulfate

Extra Examples:

Ex 3: Write the names for the following ionic compounds.

a) $CaSO_4$

c) K_2CrO_4

b) $NaNO_3$

d) Li_3PO_4

Extra Examples:

C. Multivalent Ions

- certain transition metals can form more than one type of ion, each with a different charge

Eg. Fe^{3+} Fe^{2+}

- The more commonly occurring is listed on top, thus Fe^{3+} is more common than Fe^{2+}

Copper for example forms two completely different compounds when it reacts with chlorine – one is white (1+), the other is yellow (2+).

Cu^{1+} Cl^- → CuCl (copper (I) chloride)
Cu^{2+} Cl^- → $CuCl_2$ (copper (II) chloride)

Example 1: What compound is described by Fe_2O_3?
Solution: ? 2-
Fe O

Since 3 oxygen give a charge of 6-, the charge on the iron must be 6+. Thus the charge of each of 2 iron must be 3+. Thus the compound is iron (III) oxide.

Note: The method of naming using Roman Numerals is called the Stock System.

Roman Numerals:

Ex2: Write the chemical formulas.

a) Iron (II) sulphide

b) lead(IV) oxide

c) Copper (I) Chloride

d) Copper (II) Chloride

Extra Examples:

Ex3: Write the names for the following ionic compounds.

a) NiS

b) MnF_4

c) CuF

d) Cr_2O_3

Extra Examples:

D. Hydrated Ionic Compounds

- Water molecules are loosely held within the ionic compound
- Produce water when they decompose upon heating

An example of a hydrated compound is copper (II) sulfate pentahydrate, $CuSO_4{\bullet}5H_2O$ (bluestone)

This formula indicates that 5 molecules of water are bonded within the ionic crystal for every 1 formula unit of $CuSO_4$.

Common prefixes used in naming hydrated compounds are:

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

Ex1: Write the chemical formula.

a) barium hydroxide octahydrate

b) Zinc chloride hexahydrate

c) Copper(II)sulphate pentahydrate

d) lithium hydroxide tetrahydrate

Extra Examples:

Ex2: Write the names for the following ionic compounds.

a) $MgSO_4 \cdot 5H_2O$ b) $FeCl_2 \cdot 4H_2O$

c) $Ca(NO_3)_2 \cdot 6H_2O$ d) $HgNO_3 \cdot H_2O$

Extra Examples:

<u>General Rules-Summary</u>

1. Write each ion symbol with charge: simple ions from front of periodic table; complex on back.

2. Assign the correct subscripts to each ion (subscript indicates the number of the ion preceding it).
   Total positive charge = total negative charge.

a) Switch charges to give subscripts: charge on cation becomes subscript for anion, and vice-versa.
   Use lowest whole number ratios.

OR

b) Find lowest common denominator for the charges of the 2 ions. For each ion, multiply the ion charge by a subscript number that will give the common denominator.

3. Write symbol of each ion with subscript. **Do not include ion charges in final formula.**

Note: For multivalent ions, the Roman numeral = charge on the cation, **not** the subscript.
For hydrates, prefix indicates the number of water molecules present in the ionic compound.

If more than one complex ion is present, place brackets around it, then give subscript.

Solubility of an Ionic Compound

Dissociation

< refers to the separation of ions that occurs when an **ionic compound** dissolves in water.

$$Al_2(SO_4)_{3(s)} \rightarrow 2\ Al^{3+}_{(aq)} + 3\ SO_4^{2-}{}_{(aq)}$$

< the solubility of an ionic compound will determine whether or not it will completely dissociate into ions in solution:

i) high solubility (aq) will completely dissociate into ions;
ii) low solubility (s) will not completely dissociate.

SOLUBILITY OF IONIC COMPOUNDS AT SATP - GENERALIZATIONS

Anion	Cl^- Br^- I^-	S^{2-}	OH^-	SO_4^{2-}	CO_3^{2-} PO_4^{3-} SO_3^{2-}	CH_3COO^-	NO_3^-
High Solubility (aq)	most	Group 1 NH_4^+ Group 2	Group 1 NH_4^+ Sr^{2+} Ba^{2+} Tl^+	most	Group 1 NH_4^+	most	all
Low Solubility (s)	Ag^+ Pb^{2+} Tl^+ Hg_2^{2+} (Hg^+) Cu^+	most	most	Ag^+ Pb^{2+} Ca^{2+} Ba^{2+} Sr^{2+} Ra^{2+}	most	Ag^+	none

All Group 1 compounds, including acids and ammonium compounds, are assumed to have high solubility in water.

Q. Predict which of the following ionic compounds will dissociate and write a dissociation equation. $CaSO_4$, NH_4OH, Na_2SO_4, $MnCO_3$, $Mg(NO_3)_2$

A: Check solubility using table.

MOLECULAR SUBSTANCES:

- are solids, liquids or gases at SATP
- if soluble, dissolve in water to form **colorless** aqueous solutions that **do not conduct** electricity ie. they are ***non-electrolytes***
- they contain only **nonmetal atoms**

***Molecule*:** a particle of a molecular substance that contains a fixed number of covalently-bonded nonmetal atoms

Covalent Bond: formed from the sharing of valence electrons between nonmetal atoms, which results in an electron structure that is the same as a noble gas, for each atom in the molecule

Example: **H_2** A molecule of hydrogen gas has 2 atoms of Hydrogen, each with one electron. When they bond they share a pair of electrons (one pair = one covalent bond). Since each atom now has 2 electrons, they both have the same electron structure as He (noble gas)

Diagram:

Ions DO NOT form molecules.

Molecular Substances Include Molecular Elements and Molecular Compounds

1. Molecular elements: contain only **one kind of nonmetal atom**

Type	Molecular Elements
Monatomic – one atom	Noble gases $He_{(g)}$ $Ne_{(g)}$ $Ar_{(g)}$ $Kr_{(g)}$ $Xe_{(g)}$ $Rn_{(g)}$
Diatomic – two atoms/molecule	Hydrogen, Oxygen, Nitrogen and the Halogens The "HONorable Halogens" $H_{2(g)}$ $O_{2(g)}$ $N_{2(g)}$ $F_{2(g)}$ $Cl_{2(g)}$ $Br_{2(l)}$ $I_{2(s)}$ $At_{2(s)}$
Polyatomic – more than 2 atoms/molecule	ozone = $O_{3(g)}$ Phosphorus = $P_{4(s)}$ Sulfur (Sulphur) = $S_{8(s)}$

2. Molecular Compounds-

a) Common (**to memorize**):

$H_2O_{(l)}$ = water	$CH_{4(g)}$ = methane	$CH_3OH_{(l)}$ = methanol
$H_2O_{2(l)}$ = hydrogen peroxide	$C_3H_{8(g)}$ = propane	$C_2H_5OH_{(l)}$ = ethanol
$NH_{3(g)}$ = ammonia	$C_6H_{12}O_{6(s)}$ = glucose	$C_{12}H_{22}O_{11(s)}$ = sucrose

b) Binary Molecular Compounds

- composed of 2 different kinds of nonmetals eg. CO CO_2 CCl_4 SO_3 N_2O

Writing Molecular Formulas

General Rules

1. Write each atom symbol.
2. Each prefix indicates the subscript for the nonmetal atom that precedes it (# of atoms present).
3. If no prefix is present, then there is only one atom of that nonmetal present. Monoxide = one oxygen atom present. Prefixes are the same as for hydrates.

NO BALANCING OF CHARGES OR ANYTHING ELSE. PREFIXES

Examples: Fill-in the table by writing the molecular formulas

Name	**Formula**	**Name**	**Formula**
Carbon monoxide		Trisulfur hexaoxide	
Carbon tetrachloride		Dinitrogen pentaoxide	
phosphorus pentachloride		Dnitrogen pentaoxide	
tetraphosphorus decaoxide		tetraphosphorus octaoxide	

Note: Formulas for common molecular substances **must be memorized**, as well as those for the "**HON**orable **Halogens**": $\mathbf{H_{2(g)}}$ $\mathbf{O_{2(g)}}$ $\mathbf{N_{2(g)}}$ $\mathbf{F_{2(g)}}$ $\mathbf{Cl_{2(g)}}$ $\mathbf{Br_{2(l)}}$ $\mathbf{I_{2(s)}}$ $\mathbf{At_{2(s)}}$

Naming Molecular Substances

General Rules

1. First element is named in full.
2. Second element name is shortened and given an **ide** ending.
3. Use prefixes (same as for hydrates) to indicate the number of each kind of atom.

   Prefixes:

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

4. The prefix <u>mono</u> is only used for the second element. Ex CO = Carbon monoxide. The name never starts with mono.
5. Certain Hydrogen compounds (those with H first in the formula) do not use prefixes.

   Ex. $H_2S_{(g)}$ = hydrogen sulfide, **not** dihydrogen sulfide

Examples: Fill-in the table by naming the molecular substances.

Formula	**Name**	**Formula**	**Name**
$N_2O_{(g)}$		H_2O	
$SO_{3(g)}$		H_2S	
$P_4O_{6(s)}$		NH_3	
N_3O_7		H_2O_2	

Extra Examples:

ACIDS

- Molecules that ionize in water to produce hydrogen ions, $H^{+}_{(aq)}$, ions which give acids their properties
- Most hydrogen compounds are named as acids since they form conducting solutions.
- Three exceptions are HCl(g) – hydrogen chloride, H_2S(g) – hydrogen sulfide and HCN(g) – hydrogen cyanide.

Properties of acids:

- Conduct electricity
- Turn blue litmus paper red
- Taste sour
- React with many metals to produce hydrogen gas, $H_{2(g)}$
- Have a pH value of less than 7. pH is a measure of the acidity of a substance. A pH of 7 is neutral while greater than 7 is basic and less than 7 is acidic.

PH Scale Diagram:

- Neutralize or partially neutralize bases

- General Formula: H____ $_{(aq)}$ or CH_3COO____ $_{(aq)}$
- *Note:* not all hydrogen containing compounds are acids
  Eg: NH_3 CH_4 CH_3OH C_2H_5OH

General Rules for Naming Acids:

- Name the hydrogen compound like an ionic compound, then convert the ionic name to the acid name
  hydrogen _______ide becomes **hydro______ic acid**
  hydrogen________ite becomes __________**ous acid**
  hydrogen________ate becomes __________**ic acid**

Examples: **Fill-in the table by naming the acids.**

Acid Formula	**Ionic Name**	**Acid Name**
$HCl_{(aq)}$		
$HCN_{(aq)}$		
$HNO_{2(aq)}$		
$H_2SO_{3(aq)}$		
$HNO_{3(aq)}$		
$H_2SO_{4(aq)}$		
$H_3PO_{4(aq)}$		
$CH_3COOH_{(aq)}$		

General Rules Writing Acid Formulas:

1. Translate acid name into ionic name:
   hydro___ic acid → hydrogen ___ide
   ___ous acid → hydrogen ___ite
   ___ic acid → hydrogen ___ate

2. Write chemical formulas for each ion, using rules for writing formulas for ionic compounds.

3. Hydrogen symbol is written first (cation), except for carboxylic acids (those with COO group), in which case hydrogen is placed at the end eg: CH_3COOH

4. Give the state as aqueous = (aq). We use aqueous since acids are usually in solution when used in reactions. The definition of an aqueous solution is one in which water is the solvent.

Examples: **Fill-in the table by writing the formula for the acids.**

Name	**Formula**
Hydroiodic acid	
Chlorous acid	
Chloric acid	
Boric acid	
Benzoic acid	

Extra Examples:

BASES

- Most are ionic compounds that contain the hydroxide ion, OH^-, an ion that gives bases their properties

Properties of bases:

- Conduct electricity
- Turn red litmus paper blue
- Taste bitter
- Feel slippery
- Have a pH value greater than 7
- Neutralize or partially neutralize acids

Note: Not all compounds that contain OH are bases

Eg: CH_3OH C_2H_5OH

Naming Bases: follow the general rules given for ionic compounds

Examples: NaOH ______________________________

NH_4OH ______________________________

Writing Base Formulas: follow the general rules given for ionic compounds

Examples: Lithium hydroxide ______________________________

Calcium hydroxide ________________________________

CHEMICAL REACTIONS

Types of changes in matter: Physical, chemical and nuclear

Each type of change involves changes in energy; amount increases from physical to chemical to nuclear.

1. **Physical Changes**:
   - fundamental particles remain unchanged, therefore no change in chemical formula

   E.g. Phase (state) change: $H_2O_{(s)} \rightarrow H_2O_{(l)} \rightarrow H_2O_{(g)}$

2. **Chemical Changes**:
   - involve changes in chemical bonds between atoms and/or ions
   - old bonds are broken (reactants) and new bonds are formed (products)
   - a rearrangement of atoms and/or ions occurs, therefore chemical formulas do change

   Ex $2Na_{(s)} + 2H_2O_{(l)} \rightarrow H_{2(g)} + 2NaOH_{(aq)}$

   - at least one new substance is formed, with different properties than the reactants
   - may be accompanied by changes in color, odor, state (solid precipitate or gas)
   - always accompanied by a change in energy: a net amount is absorbed or released

Physical Change	**Chemical Change**
No new or different substance is formed. The composition of the substance that undergoes the change remains unchanged	Results in the formation of at least one new substance. The particles of the new substance are different from the particles of the original substance
It is temporary change and in most cases it can be reversed by the reversal of conditions	It is permanent change and cannot be reversed by mere reversal of conditions

Physical Change	Chemical Change
No change occurs in the mass of the substances undergoing the change	Mass of the individual substances that undergo the change, always, either increases or decreases. However, the total mass of all the reactants is equal to the total mass of all the products

3. **Nuclear Change:** involves the changes within the nuclei of atoms.

CHEMICAL TESTS:

These are distinctive chemical reactions that allow you to identify an unknown substance.

1. **Oxygen test:** If a glowing splint, held in a gas, bursts into flame, then $O_{2(g)}$ is present.
2. **Carbon Dioxide:** If limewater, a clear, colorless solution of calcium hydroxide, turns cloudy (white precipitate forms), then carbon dioxide is present.
3. **Hydrogen test:** If a POP sound is heard when a burning splint is held in a gas, then $H_{2(g)}$ is present.
4. **Water:** If cobalt (II) chloride paper changes from blue to pink, then water is present.
5. **Acid:** If blue litmus paper turns red, an acid is present.
6. **Base:** If red litmus paper turns blue, a base is present.

Energy Changes in Chemical Reactions:

In a chemical reaction, energy is absorbed to break bonds of reactants and is released as new bonds form in the products.

1. **Exothermic reactions:**
   - release a net amount of energy
   - more energy is released by the products than is absorbed by the reactants

   E.g. Combustion of coal: $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + \text{Energy}$
2. **Endothermic reactions:**
   - absorb a net amount of energy
   - more energy is absorbed by the reactants than is released by the products

   E.g. Decomposition of $CaCO_3$: $CaCO_{3(s)} + \text{Energy} \rightarrow CaO_{(s)} + CO_{2(g)}$

Note: **Physical changes are also exothermic or endothermic**

E.g. melting of ice: $H_2O_{(s)} \rightarrow H_2O_{(l)}$ ***Endothermic***

E.g. freezing of water: $H_2O_{(l)} \rightarrow H_2O_{(s)}$ ***Exothermic***

Law of Conservation of Energy:

Energy is neither created nor destroyed (during any chemical or physical change), but can be converted from one form of energy to another.

e.g. photosynthesis: $6\ CO_{2(g)} + 6\ H_2O_{(g)} + \text{Energy} \rightarrow C_6H_{12}O_{6(s)} + 6\ O_{2(g)}$

Plants use sunlight (solar energy) to convert carbon dioxide and water into carbohydrates such as glucose (stored chemical energy = potential energy).

Law of Conservation of Mass (Matter):

In any chemical or physical change, mass (matter) is neither created nor destroyed.
i.e. total mass of the reactants = total mass of the products.

This leads us to believe that atoms of reactants are not changed, but simply rearranged. For example;

Methane + oxygen → water + carbon dioxide

Or symbolically, $CH_4 + O_2 \rightarrow H_2O + CO_2$

1 carbon atom	1 carbon atom
4 hydrogen atoms	2 hydrogen atoms
2 oxygen atoms	3 oxygen atoms

Note that the numbers of carbon, hydrogen and oxygen atoms are different. Have we violated the law of conservation of mass? No, we have simply not correctly written the balanced chemical equation describing the reaction.

- Demonstrated by Lavoisier (1700's), the father of modern chemistry.
- His experiments led to the emphasis on quantitative measurement, close observation and careful recording of data. All of his experiments were carried out in closed vessels.

Ex1. Lavoisier dehydrated gypsum by heating it.

$CaSO_4 \bullet 2H_2O_{(s)} \rightarrow CaSO_{4(s)} + \mathit{2}\ H_2O_{(g)}$

172.2 g → 136.2 g + ________

Ex2: Lavoisier burned mercury in air. Then he heated the product further, which decomposed back into its elements.

mercury + oxygen → mercuric oxide → mercury + oxygen

201 g + 32 g → ________ → ______ + ______

Ex3: He also burned phosphorus in air:

$P_{4(s)} + \mathit{5}\ O_{2(g)} \rightarrow P_4O_{10(s)}$

124 g + ______ → 284 g

CHEMICAL EQUATIONS

- These show the rearrangement of atoms and/or ions that takes place as a result of a chemical change of reactants into products.
- Chemical equations are a shorthand method of representing what experimental evidence indicates happens in a chemical reaction.

BALANCING CHEMICAL EQUATIONS

Balancing chemical equations involves using experimental evidence form chemical reactions. The experimental evidence indicates that:

- **Atoms are conserved**
- **Mass is conserved**
- **Energy is conserved**

A chemical equation must:

- Represent the correct chemical formula and state for each reactant and product
- Show that atoms or ions are conserved:
- Total # of each kind of atom/ion in reactants = total # of each kind of atom/ion in products

General Steps:

1. Balance atoms by using *coefficients* (in front of chemical formulas) to indicate the number of formula units or molecules of each reactant and product required.

2. Generally, begin by balancing the atom of which there is the greatest number. Find the lowest common multiple of the number of reactant and product atoms.
3. Continue progressively to balance the rest of the atoms.

Example: **Copper metal reacts with silver nitrate solution to produce silver and aqueous copper (II) nitrate.**

Note: Subscripts (s), (l) and (g) are used to indicate solid, liquid or gas and (aq) indicates an aqueous solution (in water).

Step 1: Write the word equation
Copper metal + silver nitrate → silver + copper (II) nitrate

Step 2: Write the chemical formulas
$Cu_{(s)} + AgNO_{3(aq)} \rightarrow Ag_{(s)} + Cu(NO_3)_{2(aq)}$

Step 3: Count atoms of each type for reactants and products.

1 Cu	**1 Cu**
1 Ag	**1 Ag**
1 NO_3	**2NO_3**

Note that NO_3 is treated as one complex ion.

Step 4: Balance the equation.
$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow 2Ag_{(s)} + Cu(NO_3)_{2(aq)}$

Note that balancing the nitrate ion affects the silver, which must also be balanced.

Examples: Balance the equations

1. ___ $Mg_{(s)}$ + ___ $O_{2(g)}$ → ___ $MgO_{(s)}$

2. ___$Cu_{(s)}$ + ___ $AgNO_{3(aq)}$ → ___ $Ag_{(s)}$ + ___ $Cu(NO_3)_{2(aq)}$

3. ___ $Pb(NO_3)_{2(aq)}$ + ___$KI_{(aq)}$ → ___ $PbI_{2(s)}$ + ___ $KNO_{3(aq)}$

4. ___ $NH_{3(g)}$ → ___ $N_{2(g)}$ + ___ $H_{2(g)}$

5. ___$CH_{4(g)}$ + ___$O_{2(g)}$ → ___ $CO_{2(g)}$ + ___ $H_2O_{(g)}$

6. ___ $Fe_{(s)}$ + ___ $O_{2(g)}$ → ___ $Fe_2O_{3(s)}$

7. ___ $Na_{(s)}$ + ___ $Cl_{2(g)}$ → ___ $NaCl_{(s)}$

8. __ $ASCl_{3(aq)}$ + ___ $H_2S_{(aq)}$ → ___ $As_2S_{3(s)}$ + ___ $HCl_{(aq)}$

9. __$H_2SO_{4(aq)}$ + __ $NaHCO_{3(s)}$ → _ $Na_2SO_{4(aq)}$ + __ $CO_{2(g)}$ + __ $H_2O_{(l)}$

10. ___ $C_3H_{8(g)}$ + ___ $O_{2(g)}$ → ___ $CO_{2(g)}$ + ___ $H_2O_{(g)}$

WRITING BALANCED CHEMICAL EQUATIONS

To write a balanced chemical equation from a statement or word equation:

- write the chemical formulas for all reactants and products involved (including states)
- Follow the steps outlined above for balancing equations

Examples:

Translate each of the following statements into word equations, then balanced chemical equations. Remember that ***The "HONorable Halogens" are all diatomic*****.**

1. Hydrogen and chlorine react to produce hydrogen chloride gas.

Word Equation:

Chemical Equation:

2. Solid potassium and aqueous magnesium chloride react to produce solid magnesium and aqueous potassium chloride.

Word Equation:

Chemical Equation:

3. Solid aluminum combines with oxygen gas to produce solid aluminum oxide.

Word Equation:

Chemical Equation:

4. Hydrogen peroxide decomposes (breaks down) into water and oxygen gas.

Word Equation:

Chemical Equation:

5. Zinc reacts with hydrochloric acid to produce zinc chloride solution and hydrogen gas.

Word. Equation:

Chemical Equation:

6. The combustion (burning) of ethyne gas, $C_2H_{2(g}$ in the presence of oxygen gas produces carbon dioxide gas and water vapor.

Word Equation:

Chemical Equation:

TYPES OF CHEMICAL REACTIONS

Reactions can be classified according to different types.

1. **Synthesis (*Formation, Composition)***
   - 2 elements (or 2 compounds) react to produce a single compound
   - states of reactants and products: usually all pure substances except acids (aq)

   Element + element → compound
   A + B → AB
   Where A and B are atoms and /or molecules and AB is a larger molecule.

Examples: a) magnesium reacts with oxygen from the air

b) hydrogen and oxygen react to produce water

2. **Decomposition**
   - A single compound is broken down (decomposed) into 2 or more products (elements &/or compounds)
   - states of reactants and products: usually all pure substances except acids (aq)
   - most require energy as heat, light or electricity

   Compound → two or more elements or compounds
   AB → A + B

Examples: a) mercury(II)oxide decomposes

b) water is broken down into its elements

3. **Single Displacement (*Single Replacement)***

- An element and a compound react to produce a new element and new compound
- Metal elements replace the <u>cation</u>: metal ions in ionic compounds or H^+ ions in acids or water
- Nonmetal elements replace the <u>anion</u>: nonmetal ions in ionic compounds
- States of reactants and products
    - ➢ Metal elements: all pure substances (solid except for mercury, $Hg_{(l)}$).
    - ➢ Nonmetal elements: all pure substances (solid, liquid or gas).
    - ➢ Compound reactants: usually aqueous solution (aq) or water, **$HOH_{(l)}$**
    - ➢ Compound products: if **ionic**, use solubility chart on back of periodic table
        - a) If compound is high solubility = aqueous (aq)
        - b) If compound is low solubility = solid (s)

Generally single displacement reactions follow the pattern,
Element + compound → compound + element
Z + AB → ZB + A Where Z is a metal.
Or
Element + compound → compound + element
Y + AB → AY + B Where Y is a nonmetal
Ex1: potassium + calcium iodide → calcium + potassium iodide
Z + AB → A + ZB

Ex2: Bromine + calcium iodide → iodine + calcium bromide
Y + AB → B + AY

How do we decide which element is displaced? Generally, **metals replace metals** and **nonmetals replace nonmetals**. In the example above, iodine replaced bromine (both nonmetals).

Examples:

a) mercury and silver nitrate solution react

b) zinc reacts with sulfuric acid

c) calcium reacts with water

d) Chlorine reacts with sodium bromide solution

4. **Double Displacement (*Double Replacement)***

- Usually 2 ionic compounds in aqueous solution are reacting
- Products may be one or more of:
  - Low solubility, therefore forms a precipitate (solid)- use solubility chart
  - A gas (that bubbles out of the mixture)
  - A molecular compound such as water ($HOH_{(l)}$)

compound + compound → compound + compound
AB + XY → AY + XB Where A and X are metals and B and Y are nonmetals

Example: Write the balanced chemical equation for the word equation given:
Potassium sulfate + copper (II) bromide → copper (II) sulfate + potassium bromide

Examples:

a) solutions of barium chloride and potassium carbonate react

b) solid iron(II)sulfide reacts with hydrochloric acid (one product is a gas)

5. Neutralization reaction

A neutralization reaction occurs between an acid and a base. The products of such a reaction are a salt and water.

Salt: Ionic compound that is produced by the reaction of an acid with a base.
Acid + Base → Salt + Water

An acid is added to a basic solution, the base is gradually consumed. When the entire base has reacted, the result is a neutral solution of a salt and water. The solution is neither acidic nor basic. Any additional acid will make the solution acidic.

Examples: Write the chemical equation for the following:
a) The reaction between hydrochloric acid and sodium hydroxide.

b**)** The reaction between sulfuric acid and sodium hydroxide.

c) The reaction between sodium hydroxide and sulfuric acid

Examples of neutralization reactions:

- Oven cleaner
- Baking (baking soda + acids → CO_2 bubbles which get trapped in batter causing it to rise)
- Antacids (neutralize stomach acid)
- Swimming pools
- Soda-acid fire extinguishers ($H_2SO_{4(aq)} + 2NaHCO_{3(s)} \rightarrow Na_2SO_{4(aq)} + 2CO_{2(g)} + 2H_2O_{(l)}$

6. **Combustion**

What is complete combustion?

It is the rapid reaction of a substance with oxygen to produce compounds called oxides. More commonly referred to as burning.

Generally, **Hydrocarbon + oxygen → carbon dioxide + water vapour + energy (exothermic reaction)**

$$C_xH_y + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)} + E$$

-Energy produced is in the form of heat or light.

-Fuels used in our society are mainly hydrocarbons (gas, kerosene, candles, etc.) .

-Because of the high heat involved water is produced as a gas. Also, CO_2 is produced in such large amounts it is a contributor to the greenhouse effect.

Note: In balancing hydrocarbon combustion reactions, it is easiest to balance the C and H atoms first and the oxygen last.

Examples:

a) Combustion of propane $C_3H_{8(g}$.

b) Butane, $C_4H_{10(g)}$ is burned as fuel in a lighter

c) A candle, assume $C_{25}H_{52(s)}$, combusts in the presence of oxygen

Incomplete combustion: This occurs when not enough oxygen is available. In this case the products are carbon monoxide (CO – an extremely poisonous gas), carbon (C), carbon dioxide (CO_2), and water (H_2O). Incomplete combustion does not generate as much heat energy as complete combustion.

Example: Incomplete combustion of $C_3H_{8(}$g)

Example: Incomplete combustion of $C_4H_{8(}$g)

REACTION TYPES – GENERALIZATIONS

REACTION TYPE	GENERALIZATION	STATES
Formation (Synthesis)	2 elements (or 2 cpds) → single cpd	all pure substances
Decomposition	single cpd → 2 or more products (elements &/or cpds)	all pure substances
Single Displacement	element + cpd → element + cpd	pure elements, $HOH_{(l)}$ or aqueous reactants, *s or aq products
Double Displacement	2 cpds→ 2 new cpds	aq reactants, $HOH_{(l)}$ *s or aq products
Complete Hydrocarbon Combustion	$C_xH_y + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$	

*Use the solubility chart to predict the state (aqueous or solid) of **ionic** products of displacement reactions.

NOTE:

1. Write water as **HOH** in displacement reactions and as **H_2O** in other types.
2. All metallic elements are **monatomic**. Eg. $Na_{(s)}$ $Pb_{(s)}$ & all are **solid**, except $Hg_{(l).}$
3. Some nonmetallic elements are **diatomic** ie. The "HONorable Halogens"
   $H_{2(g)}$ $O_{2(g)}$ $N_{2(g)}$ $F_{2(g)}$ $Cl_{2(g)}$ $Br_{2(l)}$ $I_{2(s)}$
4. All pure ionic compounds are **solids.**
5. Some pure molecular compounds are **gases**. Eg. NH_3 H_2S HCl and

Nonmetal oxides of C, N, and S ie. CO_2 CO SO_2 SO_3 NO NO_2 N_2O

6. In Single Displacement Reactions:
   - metal elements replace metal ions in ionic compounds or H in acids or water
   - nonmetal elements replace nonmetal ions in ionic compounds
7. In ionic compounds always write the cation (metal or ammonium ion) **first** in the chemical formula.
8. Some molecular elements: phosphorus = $P_{4(s)}$ sulfur = $S_{8(s)}$