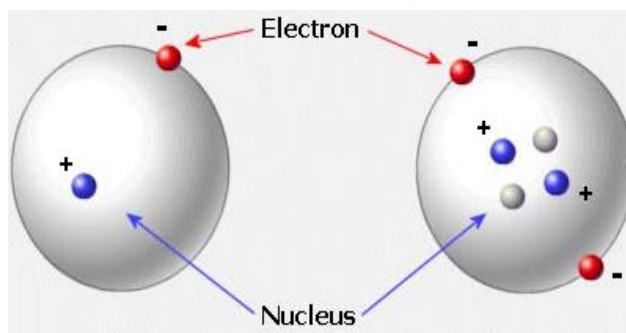
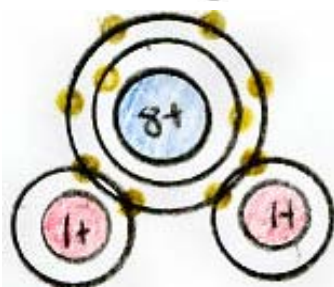
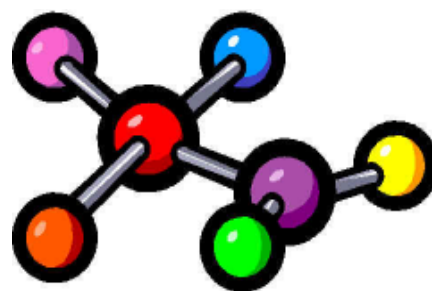
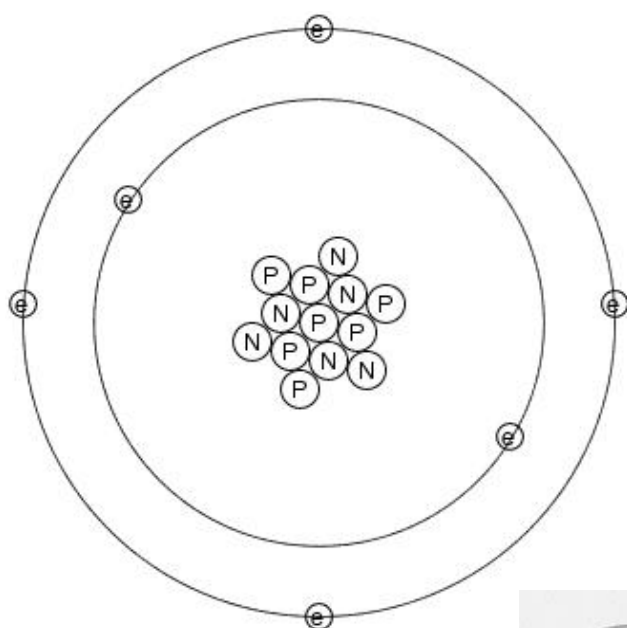


Science Module 2

# Basic Chemistry



## Teacher's Guide

# Basic Chemistry; Science Module 2

## Teacher's Guide

### Chapter One: Matter

#### A. Properties of Matter

##### Activity 1.1

Answers to all questions are yes.

The students will know everything on this page from their daily lives, though some of the words may be new to them.

It may help the students if they can see that they already know some of these ideas, but now they are just beginning to know them in science terms.

Students may have difficulty with the definition of *matter*. For instance, "stuff" and "material" may seem too simple, and not scientific. It may help to spend some time making sure that they understand this important concept.

#### B. Types of Matter

##### Activity 1.2

Elements: magnesium, copper, zinc, aluminum

Compounds: water, ethanol, sugar, stone

Mixtures: lime juice, salt water

##### Activity 1.3

Physical properties will include physical descriptions; chemical properties will include things like: combustible (burnable), etc.

#### C. States of Matter

##### Activity 1.4

It may be useful to make sure students understand the words shape (physical form), and volume (the amount of space an object takes up), before doing this activity.

Examples of different states of matter include:

solid: bottle, balloon, Eno

liquid: water,

gas: combination of Eno and water

## Activity 1.5

**sublimation** - when a solid becomes a gas

**deposition** – when a gas becomes a solid

**boiling** – This change is the rapid vaporization of a liquid, which typically occurs when a liquid is heated to its boiling point: the temperature at which the vapor pressure of the liquid is equal to the pressure exerted on the liquid by the surrounding environmental pressure. Sometimes a liquid may boil not from being heated, but when the pressure of the surrounding atmosphere is sufficiently reduced, such as the use of a vacuum pump or at high altitudes.

**boiling point** – The temperature at which the vapor pressure of the liquid equals the environmental pressure surrounding the liquid. A liquid in a vacuum has a lower boiling point than when the liquid is at atmospheric pressure. A liquid at high pressure has a higher boiling point than when the liquid is at normal atmospheric pressure. In other words, all liquids have an infinite number of boiling points.

The normal boiling point (also called the atmospheric boiling point or the atmospheric pressure boiling point) of a liquid is the special case in which the vapor pressure of the liquid equals the defined atmospheric pressure at sea level, 1 atmosphere. At that temperature, the vapor pressure of the liquid becomes sufficient to overcome atmospheric pressure and lift the liquid to form bubbles inside the bulk of the liquid. The standard boiling point is now (as of 1982) defined by IUPAC as the temperature at which boiling occurs under a pressure of 1 bar.

**vapor pressure** – All liquids and solids have a tendency to evaporate into a gaseous form, and all gases have a tendency to condense back into their original form (either liquid or solid). At any given temperature, for a particular substance, there is a pressure at which the gas of that substance is in dynamic equilibrium with its liquid or solid forms. This is the vapor pressure of that substance at that temperature.

## Activity 1.6

Some of the sugar will dissolve into the water. The sulphuric acid and water is a violently exothermic reaction. It will boil and spit. Make sure to tell the students to add the acid to the water (and not the water to the acid). Wear goggles.

## Chapter Two: Particles in Matter

### A. Atoms

#### Yes or No

1. Y
2. Y
3. N
4. N
5. N
6. N
7. N

#### The Structure of Atoms

It may be useful to draw a carbon atom and one other atom to illustrate the difference between elements. You can use the carbon atom to show its atomic number and mass and then elicit the atomic number and mass of the other atom.

Finally, draw an isotope of carbon (perhaps carbon 14) and one of the other atom, point out the difference between the two carbons and elicit the difference between the other two atoms. Repeat this with ions of the atoms.

Using a periodic chart show the students how to figure out the number of protons, neutrons and electrons for each element. Then split the students into groups, have them pick elements and repeat the process above with each other. Walk around and check their work, making sure that they correctly understand and answering any questions that come up. Having them answer each other's questions is a great confidence builder, as they see that they have the skills to help each other. Finally, have each group come up to the board and lead the class through the process.

Using the periodic chart to find the number of protons, neutrons and electrons for each element:

**protons** = the Atomic Number (whole number) is the number of protons in an atom

**electrons** = all atoms on the chart are electrically neutral, so the number of electrons are equal to the number of protons

**neutrons** = the Atomic Weight (usually number with 2 decimal places) equals the average number of particles in the nucleus. Thus the (total number of particles) – (number of protons) = (number of neutrons) or (atomic weight) – (atomic number) = (number of neutrons). Note that neutrons are whole particles (and don't exist in pieces or parts of neutrons), so the atomic weight (which is a weighted average of all the naturally occurring isotopes of an element) must be rounded to the nearest whole number.

An example using Krypton (Kr):

protons: (Atomic Number) = 36

electrons: 36

neutrons: (atomic weight) – (atomic number) or  $84 - 36 = 48$

## Exercises with Particles

These can be done in class or as homework. In either case, it may be helpful to have students write the answers on the board so that the whole class can participate in determining the correct answers.

**i.**

Element	Numbers of Protons	Numbers of Neutrons	Atomic Number	Atomic Mass*
Ca	20	20	20	40.08
Mn	25	30	25	54.94
Pd	46	60	46	106.42
Pb	82	125	82	207.2
U	92	146	92	238.03

\*Note that Atomic Masses are weighted averages and may vary from table to table.

**ii. Diagram I atoms  
Diagram II isotopes**

**iii. a. negative ion  
b. positive ion  
c. neutral atom**

**iv. a. - ion (negative ion)  
b. + ion (positive ion)  
c. + ion (positive ion)  
d. - ion (negative ion)  
e. neutral atom**

## B. Molecules

**Yes or No?**

- |      |      |      |      |
|------|------|------|------|
| 1. Y | 2. N | 3. N | 4. Y |
| 5. N | 6. N | 7. Y |      |

### Fill the gaps with words from the box.

A formula describes a molecule. Sometimes, a molecule is an element like  $F_2$ . Fluorine is called a diatomic element. An example of a molecule of a compound with 3 atoms is  $CO_2$ . An example of a formula showing three kinds of atoms is  $H_2SO_4$ .

Note: the prefix 'di-' means two, so 'diatomic' just means two atoms.

### Write the symbols for:

- |           |            |
|-----------|------------|
| 1. H      | 2. $H_2$   |
| 3. $2H_2$ | 4. $2H_2O$ |

## C. Chemical Bonds

At the heart of chemistry is physics—the study of attracting and repelling forces that link up atoms and molecules.

A **chemical bond** is the physical process responsible for the attractive interactions between atoms and molecules, and which can make diatomic and polyatomic chemical compounds stable. In general, strong chemical bonding is associated with the sharing or transfer of electrons between the participating atoms. Molecules, crystals, and diatomic gases—indeed most of the physical environment around us—are held together by chemical bonds, which dictate the structure of matter.

The explanation of these physical attractive forces is a complex area, and there are a few varieties of chemical bonds.

### Covalent and Ionic Bonds

In an ionic bond, the atoms are bound by attraction of opposite ions, whereas, in a covalent bond, atoms are bound by sharing electrons.

**Ionic bonds** are made when positively charged atoms are attracted to negatively charged atoms. These are the strongest type of bond.

**Covalent bonds** are more subtle, and occur when neighboring atoms (such as hydrogen) share electrons from within their respective halos of swirling particles. Chemists refer to both ionic and covalent bonds as "**intra-molecular**" forces.

Other important forces are called "**intermolecular**" forces—those holding different molecules together. These types of forces form the basis for liquids and solids, which are really just collections of molecules arranged in a precise pattern in space.

Intermolecular forces are also called Van der Waals forces, named after the Dutch physicist who first discovered them.

**Hydrogen bonds** are a type of Van der Waals force, and represent an important bond in biochemistry.

Bonds vary widely in their strength. Generally covalent and ionic bonds are often described as "strong", whereas hydrogen bonds and Van der Waals' bonds are generally considered to be "weak". Care should be taken because the strongest of the "weak" bonds can be stronger than the weakest of the "strong" bonds.

**Metallic bonds** - If metal atoms lose their valence (outermost) electrons, they become positive ions. If the lost electrons, which have a negative charge, stay and move about freely, the attraction of positive metal ions to all these free electrons causes a metallic bond.

In a **metallic bond**, bonding electrons are delocalized over a lattice of atoms. Because of delocalization or the free moving of electrons, it leads to the metallic properties such as conductivity, ductility and hardness.

### Exercises with Bonds

Identify the following diagrams : ionic, covalent or metallic bonds.

- a.) NaCl is ionic
- b.) Sodium metal is metallic
- c.) O<sub>2</sub> is covalent
- d.) Magnesium chloride is ionic

## Chapter Three: Chemical Reactions

This is a good time to review physical and chemical properties and physical and chemical changes from Chapter One.

### Activity 3.1

Temperature will increase, bubbles, smell. See Activity 3.4 for a more complete explanation of the reaction.

### B. Types of Chemical Reactions

#### Activity 3.2

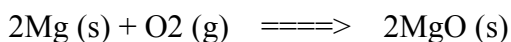
**Synthesis** is the reaction of 2 or more elements to form a compound.

General formula:  $A + B \implies C$

**Simple definition: Synthesis reactions** happen when a new substance is made from two or more other substances.

**Kinesthetic demonstration:** have two random students walk from opposite ends of the room, notice each other, say hi, and join hands. This is very funny!

An example of a synthesis reaction is:



A simple demonstration: holding the magnesium ribbon with tongs, burn the ribbon in the Bunsen flame. It should turn white. Ask students what tells them a reaction has occurred. Ask them if the ribbon looks the same as it did to begin with, and if it seems to behave differently (it's powdery now). As you can see, the magnesium oxide was formed (synthesized) from solid magnesium and oxygen gas. Have the students examine the equation for a connection to the synthesis skit.

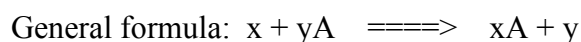
**Decomposition** is the breaking down of a large molecule into its elements or into smaller substances.



**Decomposition** is the breakdown of one substance into two or more other substances. You start off with one material and you end with several materials.

Have five or six students stand in a circle, holding hands. Have them break apart. Encourage them to sit down immediately after breaking away—large demo groups can get wacky!

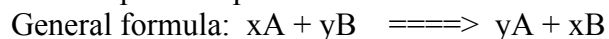
**Single replacement** is when 1 element reacts with a compound to form a new compound.



**Single replacement** reactions happen when one element replaces another element in a compound.

Two friends (a couple, if one exists in the class) walk down the aisle, and another student “takes the place” of one of the pair. Not only is this hilarious—“he stole his girl,” etc.—but it’s simple and effective.

**Double replacement** is when 2 compounds react to produce 2 new compounds. The cation of one compound replaces the cation in another compound.



In **double replacement reactions**, two pairs of elements switch partners.

Pair up two boy/girl pairs (or other types of pairs, if your classroom is progressive!) and have them switch partners, the concept is made crystal-clear. Ask questions such as, “So who was paired up to begin with?”, and “Who’s paired up now?”, to drive the point home.



Examples of reactions to give the students			
Synthesis Reactions	Decomposition Reactions	Single Replacement Reactions	Double Replacement Reactions
$2\text{Mg (s)} + \text{O}_2 \text{ (g)} \rightarrow 2\text{MgO (s)}$ As you can see, the magnesium oxide was formed (synthesized) from solid magnesium and oxygen gas.	$\text{C}_{12}\text{H}_{22}\text{O}_{11} \text{ (s)} \rightarrow 12\text{C (s)} + 11\text{H}_2\text{O (l)}$ Sucrose has decomposed into pure carbon and water	$\text{Zn (s)} + 2\text{HCl (aq)} \rightarrow \text{ZnCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$ The zinc has replaced the hydrogen in the compound.	$\text{Pb(NO}_3)_2 \text{ (aq)} + 2\text{NaI (aq)} \rightarrow 2\text{NaNO}_3 \text{ (aq)} + \text{PbI}_2 \text{ (s)}$ The lead and the sodium have 'switched partners.'
$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ The burning of hydrogen to form water	$2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2^*$ Decomposition of water by an electric current, a reaction also called electrolysis	$\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{Zn SO}_4 + \text{H}_2$ This is the replacement of hydrogen in an acid with an active metal, in this case, zinc.	$\text{NaCl} + \text{AgNO}_3 \rightarrow \text{NaNO}_3 + \text{AgCl(s)}$ The formation of solid silver chloride in an aqueous solution
$\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ The burning of carbon to form carbon dioxide	$2\text{HgO} + \text{HEAT} \rightarrow 2\text{Hg} + \text{O}_2 \text{ (g)}$ Mercury(II) oxide decomposed by heat	$\text{Cl}_2 + 2\text{NaBr} \rightarrow \text{Br}_2 + 2\text{NaCl}$ The replacement of bromide ions in a compound, by the element chlorine	$3\text{NaOH} + \text{FeBr}_3 \rightarrow 3\text{NaBr} + \text{Fe(OH)}_3 \text{ (s)}$ The formation of solid iron hydroxide in an aqueous solution
$2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$ The burning of carbon monoxide to form carbon dioxide	$2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2$	$\text{Cu(s)} + 2 \text{AgNO}_3\text{(aq)} \rightarrow \text{Cu(NO}_3)_2\text{(aq)} + 2 \text{Ag(s)}$ copper displaces silver from an aqueous solution of silver nitrate	$\text{Ca(OH)}_2 + \text{H}_2\text{SO}_4 \rightarrow 2 \text{HOH} + \text{CaSO}_4$ ( $\text{Ca}^{2+}$ replaces $2\text{H}^+$ ) note: HOH = H <sub>2</sub> O
$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$ The reaction of calcium oxide and water to form calcium hydroxide	$\text{CaCO}_3\text{(s)} \rightarrow \text{CaO(s)} + \text{CO}_2\text{(g)}$ the industrial preparation of lime (calcium oxide) involves the decomposition of calcium carbonate by heating it		$\text{AgNO}_3\text{(aq)} + \text{NaCl(aq)} \rightarrow \text{AgCl(s)} + \text{NaNO}_3\text{(aq)}$
$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ The oxidation of magnesium			
$4 \text{Fe} + 3 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3$			
$\text{CaO(s)} + \text{SO}_2\text{(g)} \rightarrow \text{CaSO}_3\text{(s)}$ calcium oxide with sulfur dioxide to form calcium sulfite			

## C. Rates of Chemical Reactions

### Activity 3.4

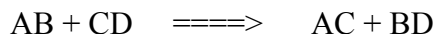
Factors that affect this set of experiments: temperature of reactants, size of reactants (surface area per volume), concentration/dilution of reactants.

Higher temperature of reactant, smaller size (increased surface area per volume), and increased concentration of reactants can speed up rate of reaction.

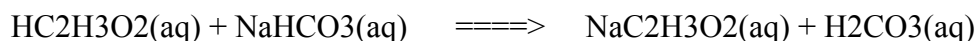
**Surface area** – The surface of something is the out side of something. An example would be the sides of a cube, the outside of a soccer ball, the skin on your hands. The surface area is the measurement of that outermost surface. For instance, a cube's surface area includes the measurements of all 6 of its sides.

**A chemical reaction occurs when acetic acid (vinegar) and sodium bicarbonate (baking soda) are mixed.**

A double displacement reaction (also called a double decomposition or metathesis reaction) has the form

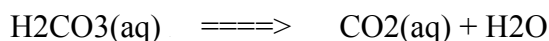


where A, B, C, and D are atoms or ions. The reaction between vinegar and baking soda can be written as a double displacement reaction if carbonic acid ( $H_2CO_3$ ) is considered a product:



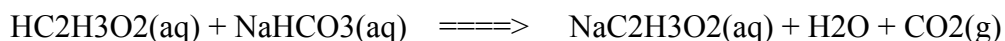
Here A, B, C, and D are  $H^+$ ,  $C_2H_3O_2^-$ ,  $Na^+$ , and  $HCO_3^-$ , respectively.

However, carbonic acid readily decomposes into carbon dioxide and water,



and the carbon dioxide can escape from the solution as a gas.

The combined equation is



which doesn't fit neatly into one category, since it is a double displacement followed by a decomposition. When you encounter a reaction that is difficult to classify, consider the possibility that it may be the net or total equation for a sequence of steps.

**A chemical reaction occurs when acetic acid (vinegar) and magnesium are mixed.**

If you mix vinegar and magnesium a reaction will occur and the temperature of the liquid will rise. The reaction is:



## Chapter Four: Acids, Bases, and Salts

### Activity 4.1

See Activity 4.2 for ideas of possible solutions

Why is water a product of the reaction of an acid and a base?

In aqueous solutions acids increase the hydrogen ion (H<sup>+</sup>) concentration. On the other hand bases increase the hydroxide ion (OH<sup>-</sup>) concentration. When an acid and a base react in an aqueous solution the H<sup>+</sup> and OH<sup>-</sup> ions combine to form water. These ions thus "neutralize" one another:

