Classifying Matter

What is Science?
Science is a way of gaining **knowledge** and **understanding** about our natural world.
Whenever we ask **why** or **how** something happens we are dealing with science.

What is Chemistry?
Chemistry is the study of **matter** and the **changes** it undergoes.

Matter
Anything that has mass and takes up space (volume)

- **Pure Substance**
  A material that has definite composition made up of atoms
  - **Element**
    A pure substance that cannot be broken down into smaller substances by chemical means
  - **Compound**
    A pure substance that is composed of two or more elements chemically combined in fixed proportions
- **Mixture**
  A combination of two or more kinds of matter, in which each component retains its own characteristics
  - **Solution (Homogeneous)**
    A mixture in which the different components are mixed so they appear to be a single substance
  - **Mechanical Mixture (Heterogeneous)** quiz/test
    A mixture in which the different components can be distinctly seen

Physical and Chemical Changes in Matter

Physical Change – a change, such as change of state, which does not alter the composition of matter.

Chemical Change – the type of change that occurs when elements and/or compounds interact with each other to form different substances with different properties.
irreversible

Five Clues that a Chemical Change has occurred:
1. A new colour or odour appears
2. Energy (heat, light, sound etc…) is given off or used up
3. Bubbles of a new gas are formed
4. A solid material (precipitate) forms in a liquid
5. The change is difficult to reverse

An Introduction to the Periodic Table

During the mid 1800's, Russian scientist **Dmitri Mendeleev** invented the modern periodic table after noticing a relationship between the **physical** and **chemical properties** of the elements. He placed the elements in order of increasing **atomic mass.** At the time approximately **58** elements had been identified.

The modern periodic table, which is comprised of over **110** elements, **92** of which are naturally occurring, is organized by **atomic number** and makes use of element **symbols** that are the same throughout the entire world.

Metals are located on the **left-hand side** and throughout the **middle** of the Periodic Table. Metals are one kind of element that have certain **properties** in common –**malleable**, **ductile**, have **lustre**, **good** conductors of **heat** and **electricity**. All metals are **solid** except for **mercury** (Hg), which is a **liquid**.

Non-metals are located on the **right hand side** of the Periodic Table. Non-metals are **brittle**, not **ductile**, not very **shiny** and **poor** conductors of **heat** and **electricity**. At room temperature non-metals may be **solids** or **gases** and one, **bromine** (Br), is a liquid.

A division line known as the "**staircase**" separates metals and non-metals. On either side of the staircase are a group of elements known as **metalloids** that show **characteristics** of both metals and non-metals.

Po Te Sb As Ge Si B

The name for each **horizontal** row of the Periodic Table is a **period**. There are **seven** periods.

The **vertical** columns in the periodic table are called **groups** and range from 1-18 (these are typically written as Roman Numerals). Some groups are given special names because they form a **family** of elements with **strong relationships.**

There are four families within the periodic table:

Group 1 – **Alkali Metals**
Group 2 – **Alkaline Earth Metals**
Group 17 – **Halogens**
Group 18 – **Noble Gases**

Periodic Chart of the Elements

alkali metals (I)

Alkaline Earth metals (II) Mmmmm

transition metals

halogens

noble gases

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca					25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn					35 Br	36 Kr
										47 Ag			50 Sn			53 I	54 Xe
										79 Au	80 Hg		82 Pb				86 Rn

inner transition metals

Atoms and Their Composition

Elements are the basic substances that make up all **matter**.

An atom is the smallest particle of an element that still retains the **identity** and **properties** of the element.

Atoms are made up of even smaller particles. These **subatomic** particles are ***protons***, ***neutrons*** and ***electrons***.

Protons and neutrons make-up the **nucleus** or core of an atom and contribute to the **mass** of an atom, while electrons are **fast moving** and occupy the **space** that surround the nucleus of the atom (**orbitals**). Electrons are so **small** and **light** that they essentially contribute no overall weight to the atom.

Subatomic Particle	Charge	Symbol	Mass (g)	Radius (m)
Electron	1-	e^-	9.02×10^{-28}	Smaller than 10^{-18}
Proton	1+	p^+	1.67×10^{-24}	10^{-15}
Neutron	0	n^0	1.67×10^{-24}	10^{-15}

Since subatomic particles are so light, chemists use a unit called an **atomic mass unit (u)** for their measurement. Both protons and neutrons have a mass of **1 u.**

Every Element has a unique:

- **Name**
- **Symbol**
  Capital letter, followed by one or two lowercase letters if present; each symbol is unique
- **Atomic number (Z)**
  equals the number of protons in the nucleus. It is inferred that the number of electrons is the same since an element is electrically neutral
- **Atomic Mass (A)**
  Equals the number of protons and neutrons in the nucleus

larger decimal # and round it!

Atomic Number **Z**

Element Name

X

atomic symbol

Atomic Mass **A**

The **mass number (M)** of an atom can be determined by **rounding** the **atomic mass** on the Periodic Table.

Examples:

17 Cl 35.45	$p^+ = 17$ $e^- = 17$ $n^\circ = 35 - 17$ $= 18 n^\circ$	7 N^{-3} 14.01	$p^+ = 7$ $e^- = 10$ $n^\circ = 14 - 7$ $= 7 n^\circ$	18 Ar 39.95	$p^+ = 18$ $e^- = 18$ $n^\circ = 40 - 18$ $= 22 n^\circ$

Mass Number:

We can use this information to calculate the number of neutrons by means of the following equation:

Number of neutrons = Mass number (M) – Atomic number (Z)

Examples:

You will notice that an element reports an **atomic mass** (a decimal number) instead of a mass number on the Periodic Table. The atomic mass represents a "**weighted average**" of all the **isotopes** for a particular atom.

Isotopes are atoms of an element that have the same number of **protons** in their nucleus, but a different number of **neutrons**.

Isotopes have very similar **chemical** properties, but they differ in **physical** properties.

Example:

"Light" Lithium

$^{6}_{3}Li$

"Heavy" Lithium

$^{7}_{3}Li$

How to Draw Atoms

Draw Bohr-Rutherford Diagrams

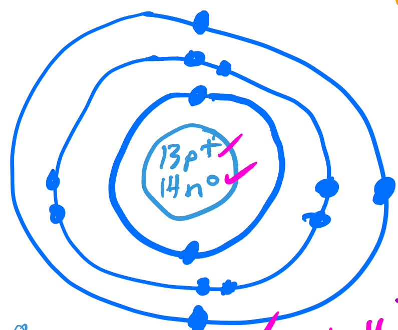
Ernest Rutherford and Niels Bohr developed the planetary model of the **atom** in 1913. In this model, the nucleus, containing the **protons** and **neutrons**, takes the central place just like the Sun takes the central place in our solar system. The electrons spin around the nucleus in orbits similar to the path of the planets around the Sun. The orbits represent the different amounts of **energy** that the **electrons** can have. Electrons in the first orbit have the **least** energy, whereas electrons in the last orbital have the **most** energy. The first orbit holds up to **2** electrons. The second and third orbits contain up to **8** electrons. As you fill the orbits, always fill the **lowest** energy orbit first, then fill up the next one and the next and so on.

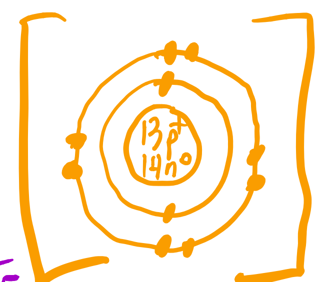
negative = anion

When you draw Bohr-Rutherford diagrams of an element, you identify the **number** of **protons** and **neutrons** in the center of the atom and place **dots** to represent the **electrons** in their orbits. Since electrons have a **negative** charge, and according the law of **electrostatics**, **oppositely** charged particles **attract** and **like** charges **repel**; you must place the first **four** electrons in the orbit as far apart as possible. For reasons beyond the scope of this course, the next **four** electrons in the orbit (if there are any) pair up with the electrons already there.

$^{27}_{13}Al$

$^{27}_{13}Al^{3+}$ isoelectric +3





cation

3 shells 2,8,3

$-3e^-$
$-8e^-$ $13p^+$
$-2e^-$ $14n^0$

Step 1: Determine the number of protons

$p^+ = 13$

This is equal to the atomic number of the element

Step 2: Determine the number of electrons

$e^- = 13$

This is equal to the number of protons.

Step 3: Determine the number of neutrons.

$n^0 = 27 - 13 = 14n^0$

Subtract the atomic number (Z) from the mass number (M) of the element. Just a reminder that the mass number is the **atomic mass** rounded to the nearest **whole number**.

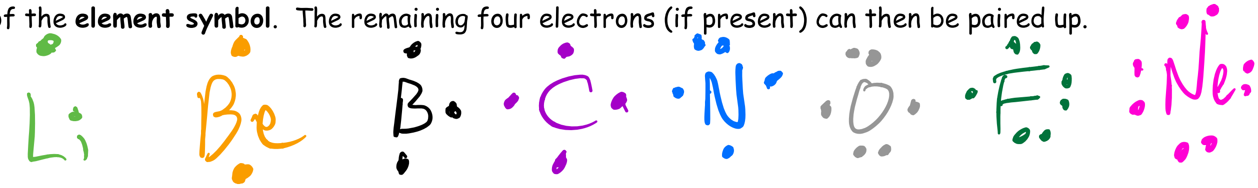
Step 4: Draw a nucleus and write in the number of protons and neutrons.

Step 5: Draw electron shells around the nucleus and fill them with the appropriate number of electrons. Always fill the inner shells to their maximum before moving to the outer shells.

Lewis (Electron) Dot Diagrams

Lewis Dot Diagrams are a short way to show the **last** energy shell (**valence** shell) for an atom. These are the electrons on the outer perimeter of an atom and generally the ones that will be involved in **bonding**.

The element **symbol** is used to represent, the **proton**, **neutrons** and all **inner electrons**. Just like when drawing B/R Diagrams, the first four valence electrons (dots) should be drawn as far apart as possible, one on each side of the **element symbol**. The remaining four electrons (if present) can then be paired up.



Classifying Chemical Compounds

A compound is a **pure substance** composed of two or more elements, chemically bonded in fixed proportions. Chemical bonds are **forces** that **attract** atoms to each other. Bonding involves the interaction between the **valence electrons** of atoms and is the driving force of **stability**.

While there are only **92** naturally occurring elements, there are **thousands** of different compounds. To help organize these compounds, chemists classify them into two main groups based on the **type** of bond that they form, and according to their **properties**.

Ionic Bond

A chemical bond between **oppositely** charged **ions** that arise from the **transfer** of **electrons**. It usually involves a **metal** and a **non-metal**.

Covalent Bond

A chemical bond in which **electrons** are **shared** by two atoms. It usually involves two **non-metals**.

Comparing Ionic and Covalent Compounds

Property	Ionic Compound	Covalent Compound
State at room temperature	Crystalline solid	Liquid, gas, solid
Melting point	High	Low
Electrical conductivity as a liquid (melted)	Yes	No
Solubility in water	Most have high solubility	Most have low solubility
Conducts electricity when dissolved in water	Strong conductor	Poor conductor

Writing Chemical Formulas

Chemical formulas are a useful way to convey information about a compound such as:

- What elements make up the compound
- The ratio or number of atoms in the compound

The chemical formula has different meanings depending on the type of **force** holding the compound together.

Covalent Compounds – Covalent compounds form **molecules**. The chemical formula of a covalent compound represents exactly **how many** of each type of **atom** are found in each individual molecule.

Example: H_2O_2 is a molecule with exactly **2 hydrogen** atoms and **2 oxygen** atoms per molecule.

Ionic Compounds – Ionic compounds form **crystals** and make a **lattice** structure. The chemical formula of an ionic compound represents a **ratio** rather than a discrete particle. Ionic compounds are always **reduced to their lowest terms**.

Example: MgO is an ionic compound that has **one** magnesium atom attached to every **one** oxygen atom in the **crystal lattice structure**.

When writing chemical formula, they are typically written such that the element found furthest to the **left** on the Periodic Table is written first.

Making Observations and Describing Matter

Observations

To notice with your **senses**. Senses may be aided by instruments such as rulers, microscopes, balances etc…

Inferences

To use **reason** and **knowledge** to make sense of your observations.

Example: The street is wet (**observation**). It rained last night (**inference**).

Observation - The fire alarm is going off.

Inference –

Observation – When a burning splint is placed in an unknown gas, the flame goes out.

Inference - the gas is $CO_{2(g)}$

Types of Observations

Qualitative Observations:

Observations **describing** the nature of something **using your senses**. For example: colour, taste, texture etc…

DOES NOT INVOLVE NUMBERS!

Quantitative Observations:

Observations describing the **amounts** or **measurements** of something. For example: how fast, how hot, how much etc…

ALWAYS INVOLVE THE USE OF NUMBERS!

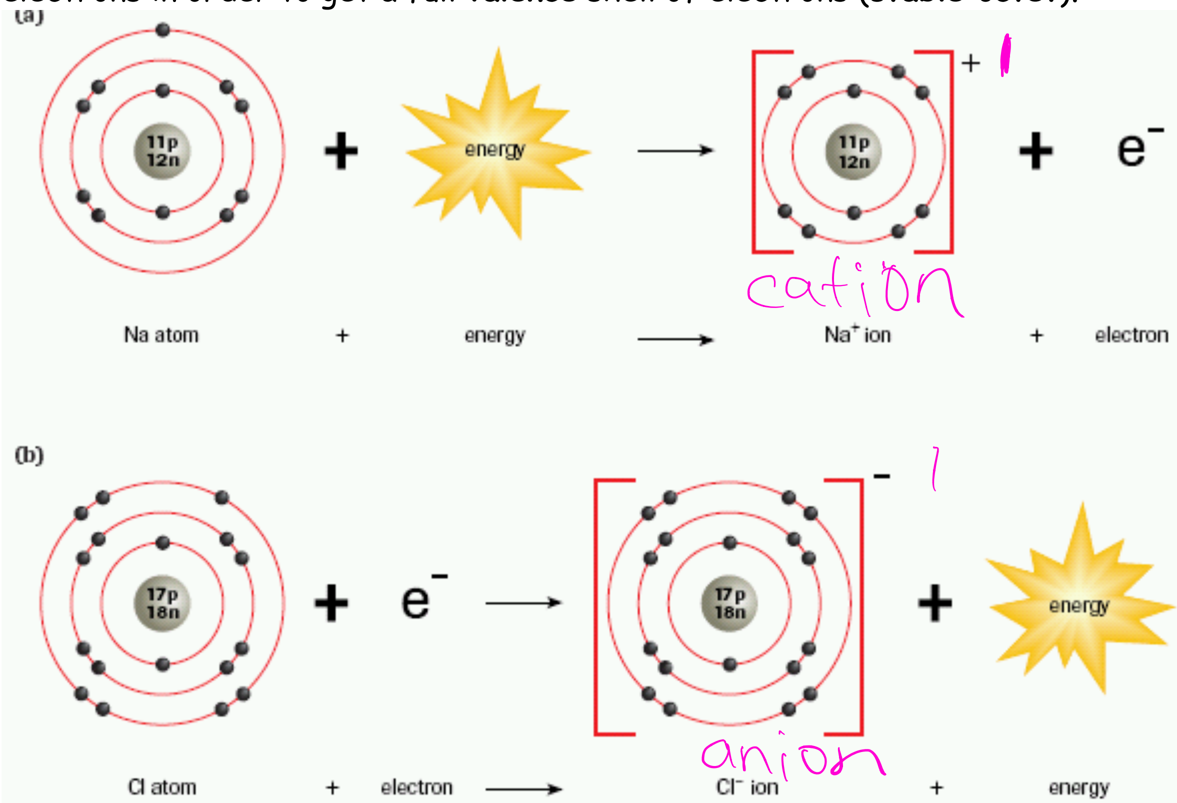
Describing matter

The properties that we can observe with our senses are called **physical properties.** The following is a list of some physical properties of matter that help us tell one thing from another.

Physical Property	Explanation or Meaning
Physical State	solid, liquid or gas
Colour	black, white, colourless, greenish-blue, yellow
Odour	odourless, spicy, sharp, flowery
Taste	sweet, sour, salty, bitter
Clarity **(transmission of light)**	1. clear (transparent) 2. cloudy (translucent) 3. opaque (no transmission)
Lustre	ability to reflect light (shiny $\rightarrow$ dull)
Form (shape)	1. crystalline (regular shape, ex. salt) 2. amorphous (irregular shape, ex. pepper)
Texture	feel - fine, coarse, smooth, gritty
Hardness	scale [1 (soft, baby powder) $\rightarrow$ 10 (very hard, diamond)]
Brittleness	ability to shatter easily (not flexible)
Malleability	Can it be hammered into a sheet?
Ductility	Can it be stretched into a wire?
Viscosity	The resistance of a liquid to flowing. Syrup is viscous, water is not.

Ionic Compounds

In order for an ionic compound to form, an atom must first become an **ion**. To do this an atom will either **gain an electron** to become an **anion** or **lose an electron** to become a **cation**. An anion has a **negative** charge and a cation has a **positive** charge. Anions and cations will **attract** to one another, forming an **ionic bond**. Atoms will **exchange** their electrons in order to get a full **valence** shell of electrons (**stable octet**).

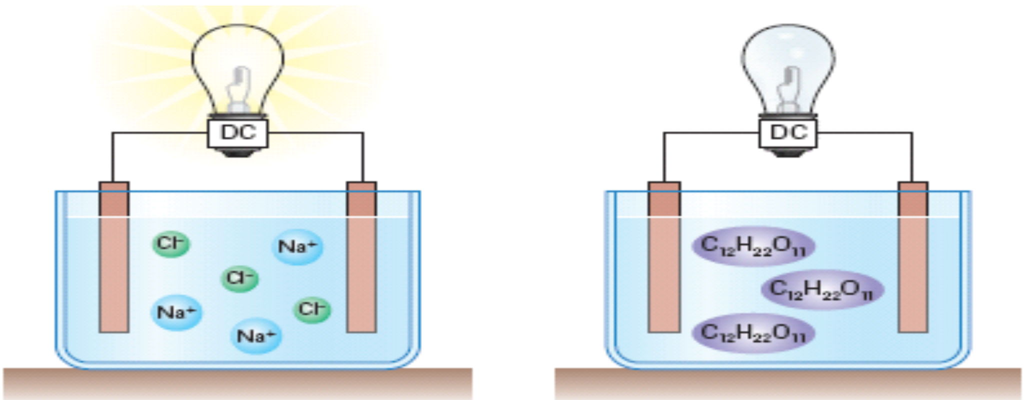


Predict the type of ion that each of the following atoms would form:

Atom	Gain or Lose Electrons	Number of Electrons	Ion Formed	Cation or Anion
Potassium				
X Magnesium	lose	2	Mg^{+2}	cation
Bromine				
Calcium				
X Nitrogen	gain	3	N^{-3}	anion
Sulphur				
X Argon	nope	Ø	Ø	—

All metals tend to form **cations** and all non-metals **anions**. Therefore, ionic compounds form when a **metal** and a **non-metal** combine. When these positive and negative particles come together they form what is called a **crystal lattice structure**; a **regular**, **repeating** pattern of ions. This is why all ionic compounds appear as **solid crystals**.

The reason that ionic compounds are capable of **conducting electricity** is because they are composed of **ions**. Electricity is the movement of **negatively charged** particles. Ionic solids are **NOT** able to conduct electricity because the ions are held in place in a **rigid crystal lattice configuration**. When **melted** or **dissolved** in water, the ions will split apart from each other (**dissociate**) and are then free to move around. A substance that can conduct electricity is termed an **electrolyte**.



Some atoms will react more intensely than others when trying to get a full outer shell of electrons.

Which of the following metals is more reactive – lithium, sodium or potassium? Can you suggest why?

K, Na, Li

Fr (most reactive)

Do you think the non-metals will follow the same pattern? For example, fluorine, chlorine and bromine? Can you suggest why?

F, Cl, Br

F (most reactive)

The following is how you can draw atoms exchanging their electrons to become ions and therefore form an ionic bond and thus becoming stable.

Bonding Atoms	EDD 1st element	EDD 2nd element	Formation of Bond (Movement of Electrons)	Ions formed	Chemical Formula
Lithium and Bromine	Li	Br	Li + Br	$[Li]^{+1}$ $[Br]^{-1}$	LiBr
Magnesium and Oxygen	Mg	O	Mg + O	$[Mg]^{+2}$ $[O]^{-2}$	MgO
Beryllium and fluorine					
Aluminium and Sulphur	Al	S	Al + S S Al S	$2[Al]^{+3}$ $3[S]^{-2}$	Al_2S_3

Covalent Compounds

Covalent compounds typically form when two or more **non-metals** bond together. During a covalent bond, valence electrons are **NOT** exchanged, but rather are **shared** between atoms. Atoms can share **one pair** of electrons, creating a **single** bond; **two pairs** of electrons, creating a **double** bond; or **three pairs** of electrons resulting is a **triple** bond. Atoms will share as many electrons as they need in order to achieve a stable octet.

Multiple Covalent Bonds				*Compound E.D.D*	*Lewis Structure*
One pair of electrons shared	→	Single bond	→		
Two pairs of electrons shared	→	Double bond	→		
Three pairs of electrons shared	→	Triple bond	→		

A unique type of interaction occurs when electrons are shared between atoms of the **same element.** There are only seven such elements that occur naturally; they are called **diatomic** molecules: $\mathbf{H_2}$, $\mathbf{O_2}$, $\mathbf{F_2}$, $\mathbf{Br_2}$, $\mathbf{I_2}$, $\mathbf{N_2}$, $\mathbf{Cl_2}$

Covalent compounds come in a variety of **states**, solid, liquid and gas, and seem to have a wide range of **properties** when compared to ionic compounds. This is due to the fact that when atoms are sharing their electrons, the sharing can occur **equally** or **unequal** and the molecules that covalent compounds form can come in a variety of **shapes**. These variations within molecules create the differences we see in covalent compounds.

Covalent compounds DO NOT typically conduct electricity when **melted** or **dissolved** in water. The atoms that make-up covalent molecules do not **break-up into ions** when they melt or boil, but rather remain as **intact molecules.** Thus, there are no **negative** charges to move around to create electricity.

The following is how you can draw atoms sharing their electrons to form covalent compounds.

Covalent Molecule	**EDD 1st element**	**EDD 2nd element**	**Compound EDD**	**Structural Diagram**	**Chemical Formula**
Oxygen and Iodine					
Phosphorous and Iodine					
Nitrogen and Fluorine					
Carbon and Bromine					
Try for a challenge: Nitrogen and Oxygen					

Chemical Reactions

A chemical reaction can be written in a number of different forms:

Chemical Equation

A description of a chemical reaction using **symbols**, not **words**, where:

- The **reactants** are written first
- The **products** are written second
- The state for each element or compound is indicated in brackets – **solid** (s), **liquid** (l), **gas** (g), **aqueous** (aq)
- Reactants and products are separated by an arrow ($\rightarrow$) – read as "**yields**"

Example:

$$\mathbf{H_2\ (g) + Cl_2\ (g) \rightarrow HCl\ (g)}$$

Word Equation

The elements and compounds that are reacting are written first followed by the products. States are included in the description.

Example:

Hydrogen gas reacts with chlorine gas to produce hydrogen chloride gas

Skeleton Equation

The Law of Conservation of Mass states that matter cannot be **created** or **destroyed**; it can only be **changed** from one form to another. Therefore the **number of atoms** in the reactants must **equal** the number of atoms in the products.

A skeleton equation is an unbalanced equation that **does not** follow the Conservation of Mass. The number of atoms on the left side (reactants) of the chemical equation **does not** equal the number of atoms on the right side (products).

Example:

$$H_2\ (g) + Cl_2\ (g) \rightarrow HCl\ (g)$$

On the reactant side there is a total of **4** atoms (**2** hydrogen and **2** chlorine)
On the product side there is a total of **2** atoms (**1** hydrogen and **1** chlorine)

Balanced Chemical Equation

A balanced chemical equation is an equation that follows the Law of Conservation of Mass. The number of atoms on the reactant side equals the atoms on the product side. In most chemical equations, numbers placed in front of the elements or compounds (**coefficients**) are required to balance the equation.

Example:

$$1H_2\ (g) + 1Cl_2\ (g) \rightarrow 2HCl\ (g)$$

On the reactant side there is a total of **4** atoms (**2** hydrogen and **2** chlorine)
On the product side there is a total of **4** atoms (**2** hydrogen and **2** chlorine)

When there is a coefficient of "**1**", it is typically not written: $H_2\ (g) + Cl_2\ (g) \rightarrow 2HCl\ (g)$

Balancing Equations

All chemical equations must be balanced so that they are consistent with the Law of Conservation of Mass.

Here are some suggestions for balancing equations:

1. When balancing equations, always start with the "ugliest" molecule first (polyatomics).

2. To balance, place the desired number (coefficient) in front of the element or compound. Never split-up a compound and never change the subscripts in the chemical formula.

3. It is often useful to balance the diatomic molecules, if they are present, last.

4. Creating a chart to keep track of the type and number of each atom on the reactant and product side of the equation can make balancing easier.

5. Make sure to always recheck the final balanced equation.

Examples:

____ Mg (s) + ____ O_2 (g) → ____ MgO (s)

Atoms	Reactants	Products
Mg		
O		

____ H_2 (g) + ____ O_2 (g) → ____ H_2O (g)

Atoms	Reactants	Products
H		
O		

____ Fe (s) + ____ O_2 (g) → ____ Fe_2O_3 (s)

Atoms	Reactants	Products
Fe		
O		

____ Al_2O_3 (s) + ____ H_2 (g) → ____ H_2O (l) + ____ Al (s)

Atoms	Reactants	Products
Al		
O		
H		

____ $Pb(NO_3)_2$ (aq) + ____ BF_3 (s) → ____ $B(NO_3)_3$ (aq) + ____ PbF_2 (s)

Atoms	Reactants	Products
Pb		
NO_3		
B		
F		

Sometimes to balance an equation, fractions must be used. Fractions are not to be left in the final balanced equation, as it impossible to have part of an atom. To get rid of the fraction, multiply every element or compound in the equation by the denominator of the fraction (i.e. If you use $\frac{1}{2}$ as a coefficient, then multiply by 2).

____ NH_3 (l) + ____ O_2 (g) → ____ NO_2 (g) + ____ H_2O(s)

Atoms	Reactants	Products
N		
H		
O		

____ FeS_2 (s) + ____ O_2 (g) → ____ Fe_2O_3 (s) + ____ SO_2 (g)

Atoms	Reactants	Products
Fe		
S		
O		

Balancing chemical equations becomes increasing more difficult when you are given the reaction as a word equation. To balance the equation, you must first convert the elements and/or compounds into their correct chemical formula. Even the slightest mistake will make you equation incorrect and could possibly create an equation that is impossible to balance. Be careful, and make sure to always check your work.

Write out a balanced chemical equation for the following:

Oxygen gas reacts with solid aluminum sulfide to produce solid aluminum oxide and sulfur dioxide gas.

Balancing Word Equations

Write the appropriate formulas and symbols below the word equation and then balance each reaction.

1. dicarbon dihydride gas reacts with oxygen gas to produce carbon dioxide gas and liquid dihydrogen monoxide

2. hydrogen iodide gas and aqueous sulfuric acid (hydrogen sulfate) react to produce aqueous hydrogen sulfide, iodine gas and liquid dihydrogen monoxide

3. Aqueous potassium sulfate reacts with aqueous barium nitrate to yield aqueous barium sulfate and aqueous potassium nitrate

Types of Chemical Reactions

It is important to be able to classify chemical reactions as it enables scientists to predict possible products or outcomes. For example, think of appropriate storage of chemicals…

Why are some chemicals stored in dark containers?

Why are some chemicals stored in glass jars?

Why is it inappropriate to store propane tanks in areas that get very hot?

Below are 4 major categories of chemical reactions:

1. Synthesis

A synthesis reaction occurs when 2 or more **elements** combine to form a new **molecule** or **compound**.

The general equation for a synthesis reaction is: $\mathbf{A + B \rightarrow AB}$

Specific types of synthesis reactions:

a) **Metals** react with **oxygen** to produce a **metal oxide**

b) A **non-metal** reacts with **oxygen** to produce a **non-metal oxide**

c) A **metal** and **non-metal** combine to form a **binary ionic compound**

d) **Non-metallic oxides** react with **water** to produce an **acid**

e) **Metallic oxides** react with **water** to produce a **base**

2. Decomposition

A decomposition reaction is the reverse to a synthesis reaction, a compound **breaks down** into **elements** or other **compounds**

The general equation for a decomposition reaction is: $\mathbf{AB \rightarrow A + B}$

Example:

Typically, some form of **energy** or type of **catalyst** is needed to initiate a decomposition reaction.

A catalyst is a substance that controls the **rate** of a reaction, without being **used-up** during the reaction or affecting the overall **products**.

3. Single Displacement Reaction

A single Displacement reaction occurs when one **element** in a compound is **displaced/replaced** by another **element**. This can occur in 2 ways, a **metal** can replace a **metal** or a **non-metal** can replace a **non-metal**

The general equation for a single displacement reaction is: $\mathbf{A + BC \rightarrow AC + B}$ (if A is a metal) **or** $\mathbf{X + YZ \rightarrow YX + Z}$ (if X is a non-metal).

Examples: a) $\mathbf{Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe}$

b) $\mathbf{Cl_2 + CaBr_2 \rightarrow CaCl_2 + Br_2}$

c) $\mathbf{Cu + AgNO_3 \rightarrow CuNO_3 + Ag}$

How do you know that a single displacement reaction can occur or do they always occur?

For example, explain why the two above reactions occur but the following reaction does not?

$$\mathbf{Pb + Fe_2O_3 \rightarrow \text{No Reaction}}$$

In order to determine if an element will displace another element in a single displacement reaction you must refer to the **Activity Series of Metals**:
If one element is **above** another element in the compound, it can be **bumped out** and a single displacement reaction will occur.
Non-metals, typically **halogens** are involved in Single Displacement Reactions. To determine who can bump out whom, you must refer to the **Activity Series for Halogens**.

Predict if the following reactions will occur and what the products are:

Fluorine
Chlorine $I_2 + NaCl \rightarrow$ ______________________________
Bromine
Iodine $F_2 + KBr \rightarrow$ ______________________________

4. Double Displacement Reactions

A double displacement reaction occurs when there is an **exchange** of **cations** between two **ionic** compounds.

The general equation for a double displacement reaction is:

$$\mathbf{AB + CD \rightarrow AD + CB}$$

In the general equation above, A and C are **cations** (written first) and B and D are **anions.**

How do you know that a double displacement reaction can occur or will they always occur?

Evidence that a double displacement reaction will/has occurred:

A) A <u>solid</u> precipitate (ppt) forms
B) A gas is produced, bubbles form
C) Water (H_2O) is formed

Example: $NaCl + AgNO_3 \rightarrow$ ________________________________

Example: $Na_2CO_3 + HCl \rightarrow$ ________________________________

Example: $H_3PO_4 + Ca(OH)_2 \rightarrow$ ________________________________

Water is evidence of an **acid/base** reaction (**neutralization**), which is a type of double displacement reaction. Since water is a clear, colourless, liquid, it typically cannot be seen by looking at the reaction. To determine if water is present, it has to be tested using **indicators** or **pH values**.

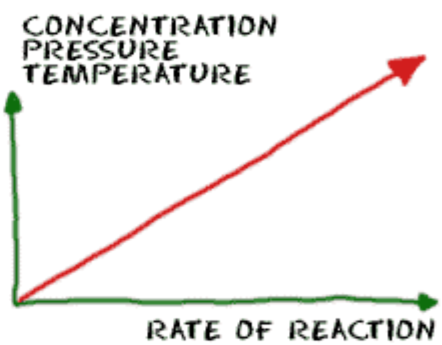
Rates of Reactions (and Energy Changes - 2DX)

Rates of Reactions

The rate of reactions is defined as the **time it takes for a given product to form or for a given amount of reactants to be used up**.

Rates of reactions can be explained using the **collision theory**. The collision theory states that the **more collisions** that occur between atoms or molecules, the **more likely a reaction** will happen. If there are a higher number of collisions in a system, more combinations of molecules will occur. The reaction will go faster, and the rate of that reaction will be higher.

Reactions happen, no matter what. Atoms are always combining or compounds breaking down. The reactions happen over and over but not always at the same speed. A few things affect the overall speed of the reaction and the number of collisions that can occur.



Concentration: If there is **more of a substance** in a system, **there is a greater chance that molecules will collide** and speed up the rate of the reaction. If there is **less of something**, there will be **fewer collisions** and the reaction will probably happen at a slower speed.

Temperature: When you **raise the temperature** of a system, the molecules bounce around a lot more (an **increase in thermal energy**). When they bounce around more, they are **more likely to collide**. That fact means they are also more likely to combine. When you lower the temperature, the molecules are slower and collide less. That temperature drop lowers the rate of the reaction.

Pressure: Pressure affects the rate of reaction, especially when you look at gases. When you **increase the pressure**, the molecules have **less space in which they can move**. That greater concentration of molecules increases the number of collisions. When you **decrease the pressure** atoms and/or molecules **spread out** and don't hit each other as often. The lower pressure decreases the rate of reaction.

Surface Area: When you **increase the area in which reactants can come into contact with each other**, you are increasing the number of atoms/molecules that are able to collide. The more collisions that occur, the greater the opportunity of a reaction occuring.

An example of this is can be seen when comparing a pack of sugar versus a sugar cube placed in water. A pack of sugar provides a greater surface area, as every sugar crystal will be in contact with the water. With a sugar cube, only the outer layer of sugar is in contact with the water and therefore capable of reacting.

Catalyst: A catalyst is defined as **a substance that controls the rate of a reaction without being used in the reaction itself**. Catalysts lower the energy required (**activation energy**) required to break the bonds that hold substances together.

Examples of catalysts include enzymes (biological systems), palladium (catalytic converters) and even light (hydrogen peroxide).

Energy Changes and Chemical Reactions (2DX content only)

All chemical reactions involve the **input and release of energy**. Often thermal energy is involved, but can the energy can also come in the form of light, electricity and sound.

You can classify reactions on the basis of whether they release or absorb more energy. **Energy releasing** reactions are called **exothermic**. Examples include the burning of fossil fuels and the rusting of iron.

Some reactions involve the addition of large amounts of energy to cause a chemical change (**large activation energy**). **Energy-absorbing** reactions are called **endothermic**. Cooking food, ice packs and electrolysis are all examples of endothermic reactions.

Identify the following as exothermic or endothermic:

Ice melting - ___________________________
A match burning - ___________________________
Frying an egg - ___________________________
Mixing acids with water will cause a rise in temperature - ________________________
Hydrogen gas and chlorine gas will explode when exposed to UV light - _______________

Acids and Bases

An **acid** is a substance that produces **hydrogen ions** in solution, $\mathbf{H^+}_{(aq)}$. For example:

i) When hydrochloric acid, HCl is placed in solution it dissociates (ionizes) into:
H^+ and Cl^-

ii) When sulfuric acid, H_2SO_4 is placed in water it dissociates (ionizes) into:
H^+ and SO_4^{-2}

A **base** is a substance that produces **hydroxide ions** in solution, $\mathbf{OH^-}_{(aq)}$.
For example:

i) When sodium hydroxide, NaOH is placed in solution it dissociates (ionizes) into: **Na^+ and OH^-**

ii) When calcium hydroxide, $Ca(OH)_2$ is placed in solution it dissociates (ionizes) into: **Ca^{+2} and OH^-**

Acids and bases have **characteristic properties** that are summarized in the table below:

Acids	Bases
Taste sour	Taste bitter
Has no characteristic feel	Feels slippery
Conducts electricity	Conducts electricity
Keeps red litmus red	Turns red litmus blue
Turns blue litmus red	Keeps blue litmus blue
Turns bromothymol blue yellow/green	Bromothymol blue remains blue
Keeps phenolphthalein clear	Turns phenolphthalein pink
Reacts with active metals to produce hydrogen gas (burning splint test)	Does not react with metals
Reacts with sodium carbonate to produce carbon dioxide (limewater test)	Does not react with sodium carbonate
Does not react with ammonium chloride	Reacts with ammonium chloride to produce ammonia (waft for odour)

Indicators

Most solutions of acids or bases are **clear** and **colourless**. Therefore they cannot be distinguished from ordinary water by appearance alone. The simplest way to distinguish them from water is to use an **indicator**. An indicator is a substance that produces a **change in colour** as the concentration of $\mathbf{H^+}$ and $\mathbf{OH^-}$ changes.

Indicators can be made from **natural products** such as flowers, fruit and vegetables. There are also a number of **synthetic indicators**. Theses are more common as they tend to last longer and can be produced in large quantities.

Concentration of Acids and Bases (pH)

Concentration is defined as the amount of **solute** per quantity of **solvent**. The concentration of a product can easily be altered by diluting with **more solvent** or the addition of **more solute**. Water is the universal solvent.

When you determine the concentration of hydrogen ions in solution (amount of H+ ions/ total solution volume) you are determining the **pH** of that particular solution. pH stands for, "**the power of hydrogen**". The pH of a substance can be determined a number of different ways, such as with the use of pH paper, an electronic pH meter or mathematically. **The pH scale ranges from 0-14**.

Acids have a pH **less than 7**
Bases have a pH greater than **7**
Neutral substances have a pH **equal to 7**

While the pH scale ranges from 0 to 14 and each pH unit represents a factor of 10.

A change in pH from 3 to 8 is a(n) ______________ increase/decrease in $[H^+]$
A change in pH from 11 to 2 is a(n) ______________ increase/decrease in $[H^+]$

Strength of Acids and Bases

Strong acid – an acid that dissociates completely into ions in water.

Example: $HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

When hydrogen chloride molecules enter an aqueous solution, 100% of the hydrogen chloride molecules dissociate. As a result the solution contains the same percent of H^+ ions and Cl^- ions: 100%

Weak acid – an acid that dissociates very slightly in a water solution.

Example: $CH_3COOH_{(aq)} \Leftrightarrow H^{+}_{(aq)} + CH_3COO^{-}_{(aq)} + CH_3COOH_{(aq)}$

On average, only about 1% of the acetic acid molecules dissociate at any given moment.

Notice that the arrow used in the dissociation of a weak acid points in both directions. This indicates that the reaction is ***reversible***. The products of the reaction will also react to produce the original reactants.

Strong base – a base that dissociates completely into ions in water.

Examples: NaOH, $Mg(OH)_2$

Weak base – a base that dissociates very slightly in a water solution.

Example: NH_3

Neutralization Reactions

Neutralization occurs when **hydroxide ions** (base) and **hydrogen ions** (acid) are mixed to make **water** and a **salt**. Neutralization reactions are types of **double displacement** reactions. The general word equation for a neutralization is:

Examples:

1. *Given the full equation in words:*

Aqueous solutions of hydrobromic acid and beryllium hydroxide undergo a neutralization reaction to produce liquid water and aqueous beryllium bromide.

2. *Given the partial equation in words or in these cases, in chemical formulae, you can complete the following equations:*

____ $H_2SO_{4\,(aq)}$ + ____ $LiOH_{(aq)} \rightarrow$

____ $Ca(OH)_{2\,(aq)}$ + ____ $H_3PO_{4\,(aq)} \rightarrow$

3. *Working backwards from the examples above, you can determine which acid and base would react together to produce the following salts:*

i) KNO_3

ii) $MgCO_3$

Elements and Oxides

An oxide is any element chemically combined with oxygen. How does the element's position in the periodic table affect the ability of the oxide to form an acid or a base?

Reactions of Metals

Review:

- Metals are found on the **left** side of the staircase
- Metals are generally shiny, ductile, malleable, good conductors of electricity and heat, and **solid** at room temperature (except **Mercury**)

There are certain patterns of chemical behavior that metals follow:

- Form **metal oxides** when they react in oxygen
- Metal oxides are always **solids**
- Metal oxides form **bases** when they react with water

Since they form **bases**, they can be called **basic** oxides or **basic anhydrides**.

For example:

Potassium burns in oxygen to produce potassium oxide. The balanced chemical equation representing this statement is:

$$4\ K + O_2 \rightarrow 2\ K_2O$$

When the potassium oxide reacts with water the product is potassium hydroxide. The balanced chemical equation representing this statement is:

$$K_2O + H_2O \rightarrow 2\ KOH\ (aq)$$

Potassium hydroxide is used in **liquid fertilizer**, **cosmetics, paint removers**, and **making soap**.

Reactions of Non-Metals

Review:

- Non-metals are found to the right of the staircase
- Non-metals are usually brittle, dull, poor conductors of heat and electricity, and have a variety of states at **room temperature**

Non-metals also follow certain patterns of chemical behavior, such as:

- Form **non-metal** oxides when they react in **oxygen**
- Non-metal oxides are often **liquid** or **gases**
- When non-metal oxides react with water they form **acids**

Since they form **acids** they can also be called **acidic** oxides.

For example:
Nitrogen reacts with oxygen to form nitrogen dioxide. The balanced equation representing this statement is:

$$N_2 + O_2 \rightarrow 2\ NO_2$$

When the nitrogen dioxide is reacted with water, the product is nitric acid. The balanced equation representing this statement is:

$$3\ NO_2 + H_2O \rightarrow 2\ HNO_3\ (aq) + NO$$

Nitric acid contributes to our **air pollution** and is used in many **industrial reactions.**