ELEMENTS OF CHEMICAL CHANGE I

1. CHEMICAL REACTIONS

- As a provider of health care, you will not be required in most cases, to write and balance chemical equations.
- You will, however, be using and/or seeing the effects of chemical reactions on a daily basis.
- Chemical reactions are frequently used to explain various concepts of pharmacology and physiology.
- Consider drugs.

- All drugs are chemicals and any pharmacological reference you consult will refer to the chemical changes drugs undergo in the body.
- Consequently, it is essential that you have a basic knowledge of what a chemical reaction involves and how that chemical reaction can be expressed as a chemical equation.

a. Definite Composition.

- When atoms combine, they do so in definite ratios of intact atoms to produce compounds with definite composition.
- Note that this combination is by number of atoms, not by weights of atoms.
- What the individual atoms happen to weigh is not important.
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a. Definite Composition.

- Atoms do not know what they weigh.
- When they do interact and combine, it is always as whole particles, and the particle-to-particle or atom-to-atom ratio can always be expressed in simple, whole numbers.
- Chemical changes do not split atoms into fractional pieces.

- This is the reason we are able to write a formula such as HCl for the compound hydrochloric acid.
- Hydrochloric acid is always formed from one atom of hydrogen and one atom of chlorine.
- Since a chemical reaction is merely a change in matter, and matter consists of atoms or molecules.

b. Chemical Equations.

- In discussing a chemical reaction, it would be very cumbersome to write it out in the same manner as we state it verbally.
- To get around this problem, chemists have developed chemical equations.
- Chemical equations are abbreviated ways of writing chemical reactions.
- They save much writing and effort and give at least as much information as a verbally stated reaction.

Chemical equations show:

(1) The kinds of atoms or molecules reacting.

(2) The products formed.

(3) The number of atoms entering the reaction.

(4)

Chemical equations show:

(4) The number of molecules formed in the product.

(5) The proportion in which the substances react to give definite products.

c. Chemical Symbols.

- In writing chemical equations, we use a number of symbols.
- The most common symbols are shown below with their meanings

SYMBOL	MEANING
Δ	Heat (a form of energy)
$\rightarrow$	"yields," indicates direction of reaction
$\uparrow$	given off as a gas
$\downarrow$	given off as a precipitate

d. Types of Reactions.

There are <u>four types of chemical reactions</u>, which are possible:

- Combination reactions,
- Decomposition reactions,
- Single replacement reactions, and
- Double replacement reactions.

(1) Combination reactions.

A combination reaction can be represented by the chemical equation A + B --> AB (one atom of A plus one atom of B yield one molecule of AB).

A specific example of this type of reaction is the combination of a metal with oxygen to yield a metallic oxide.

$$2\ Mg + O_2 \text{ --> } 2\ MgO$$

This equation tells us that two atoms of magnesium and one molecule of oxygen react to form two molecules of magnesium oxide.

(2) Decomposition reactions.

The general equation representing decomposition reactions is $AB \rightarrow A + B$.

Here is a good example:

$$CaCO_3 \xrightarrow[\Delta]{} CaO + CO_2 \uparrow$$

This equation tells us that calcium carbonate will yield calcium oxide and carbon dioxide.
The Δ also tells us that this reaction occurs when heat is applied to calcium carbonate.
The $\uparrow$ indicates that the carbon dioxide is given off as a gas.

(3) Single replacement reactions.
The general equation for a single replacement reaction is
$A + BC \rightarrow AC + B$.
An example is:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

This equation tells us that one atom of zinc and one molecule of cupric sulfate yield one molecule of zinc sulfate and one atom of copper.

(4) Double replacement reactions.

The most commonly occurring reaction is the double replacement reaction. The general equation for this reaction is

$AB + CD \rightarrow AD + CB$.

Double replacement reactions can be further subdivided into several classes. The most common of these classes are the precipitation reaction, the acid–base reaction, and the oxidation–reduction reaction.

An example of the precipitation reaction is:

$$BaCl_2 + Na_2SO_4 \rightarrow 2\ NaCl + BaSO_4$$

This equation tells us that one molecule of barium chloride reacts with one molecule of sodium sulfate to yield two molecules of sodium chloride and one molecule of barium sulfate as a precipitate.

2. WRITING CHEMICAL EQUATIONS

- One general rule that must be kept in mind is that there will always be the same number and kinds of atoms in the products of a reaction as in the reactants.

- This is because matter can neither be created nor destroyed in a chemical reaction and atoms always combine in certain proportions.

- When given a written verbal description of a chemical reaction, the following steps are used to write the equation for the reaction.

a. Write the symbols for all elements involved.

b.

- Write the correct formulas for any compounds and check for diatomic molecules.

(Some elements never exist as single atoms but only as diatomic molecules. These elements can be identified from their names, which end in –gen or –ine. The common diatomic molecules are hydrogen (H_2), nitrogen (N_2), oxygen (O_2), chlorine (Cl_2), fluorine (F_2), and bromine (Br_2).)

c.

- Balance the equation by placing coefficients where appropriate.
- Remember that there must be equal numbers of atoms of each kind on both sides of the equation.
- In this step, the subscripts that were used in writing the correct formulas cannot be changed.

3. EXAMPLE

- For application of these steps, consider this description of a reaction.
- Calcium metal and water react to yield calcium hydroxide and hydrogen gas.

a. Write the symbols for all elements involved.

Ca, O, H

b. Write the correct formulas for any compounds and check for diatomic molecules.

$$Ca + H_2O \rightarrow Ca(OH)_2 + H_2 \uparrow$$

c. Balance the equation by placing coefficients where appropriate. Look at the number of atoms of each element in the products and reactants.

REACTANTS	PRODUCTS
1 Ca	1 Ca
1 O	2 O
2 H	4 H

$$\_\_\_\_ Ca + \_\_\_\_ H_2O \rightarrow \_\_\_\_ Ca(OH)_2 + \_\_\_\_ H_2\uparrow$$

4. EQUILIBRIUM REACTIONS

- We have implied that all reactions only go in the direction of the products, but this is not always the case.
- Sometimes as products are formed, they react with one another or decompose to form the reactants.
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- Thus, the reaction is going in both directions at the same time, and if allowed to continue indefinitely, would result in a constant amount of products and reactants.

- Reactions that go in both directions are called equilibrium reactions, and when the rate of formation of product is the same as the rate of formation of reactant, they are said to be in equilibrium.

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- In writing an equation, we indicate equilibrium by drawing arrows pointing in opposite directions

Ex) $Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{-2}$

Sodium carbonate in solution dissociates into sodium ions and carbonate ions. Some of the ions come back together to form sodium carbonate. Thus, an equilibrium is established.

5. EXTERNAL CONDITIONS AFFECTING CHEMICAL REACTIONS

- External conditions that affect reactions are usually types of energy that are put into a reaction, such as heat or light.
- Chemical reactions are always accompanied by an energy change.

- Either energy is released or it is acquired.
- When the amount of energy is changed, so is the amount of matter.
- This is called the Law of Conservation of Matter and Energy.
- However, ordinary chemical reactions involve such small matter changes that they go undetected and may be ignored.

a. Heat.

- Generally, heat is the form of energy we are most concerns us most.
- It may affect a reaction in one of two ways.

(1) Exothermic reactions.

- If a reaction gives off heat, it is called an exothermic reaction.
- External heat, if supplied to this type of reaction, will slow down the rate of reaction.

(2) Endothermic reactions.

- If a reaction takes in heat, it is an endothermic reaction.
- If heat is added to an endothermic reaction, the rate of reaction will increase.
- This may be of value in the preparation of medicinal products.

b. Light.

- Light is a form of energy that may cause <u>many chemicals to decompose</u>.
- For this reason, it is necessary to protect some drugs from contact with light by placing them in dark-colored or opaque containers.
- These containers prevent most or all of the outside light from coming into contact with the drug.

6. REACTING QUANTITIES

- It has already been emphasized that <u>all reactions occur on an atom–to–atom level.</u>
- This presents a small problem to us, since we cannot hold an atom in our hand, or count out a specific number of atoms to put into a reaction.
- How then do we measure amounts of material that will react together?

- Chemists long ago solved this problem by learning how to count particles indirectly.
- They did this by measuring samples of the chemicals in particular ratios by their weights.
- To understand the means of doing this, we need to expand our concept of atomic weight to compounds in the form of the formula (or molecular) weight.

a. Milligram Formula (Milligram Molecular) Weight.

- When atoms combine to form compounds, the atomic nuclei are not affected.
- There is no net loss of weight.
- Regardless of whether the particle formed is a molecule or an ion group, it will have a formula and a formula weight.
- The formula weight of a compound is the sum of the atomic weights of all the atoms that appear in its chemical formula.

Consider, for example, carbon dioxide:

Atoms: $C + O + O = CO_2$ (molecule)

Atomic weights: $12 + 16 + 16 = 44$ (formula weight)

- While we have arrived at a formula weight which is in terms of atomic mass units, it is much more useful to express it in terms of milligrams.
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- This is known as the milligram formula weight.
- For the example above, CO2, the milligram formula weight is 44 mg.
- This is a quantity that we can measure and see, and thus can easily work with.
- It also represents a reacting unit of the compound.

b. Molarity.

- A molar solution, or a one molar (1M) solution, consists of one-gram molecular weight (GMW) of solute dissolved in enough water to make 1 liter of finished solution.
- Molarity, then, is the number of GMWs dissolved in enough water to make a finished solution of 1000 ml.
- Molar solutions may have as a solute a solid, a liquid, or a gas.

(1) Calculating the gram molecular weight.

- One-gram molecular weight of a substance is its molecular weight expressed in grams.
- Thus, a GMW of NaOH would be 40 grams, where the atomic weights are as follows: Na = 23, O = 16, and H = 1.
- Thus, 0.5 GMW of NaOH would be 20 grams, and so forth.
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- A mole is one-gram molecular weight of a substance.
- Thus, a mole of NaOH is 40 grams of NaOH; a half-mole (.5 mole) is 20 grams; two moles of NaOH are 80 grams, and so on.

(2) Calculating the molarity of a solution.

- To find the molarity of a solution, we divide the number of gram molecular weights of solute by the number of liters of total solution.
- The formula may be written:

$$\text{Molarity} = \frac{\text{no. of GMWs of solute}}{\text{no. of liters of solution}}$$

- Since many problems are stated in terms of the weight of solute and require you to determine the number of gram molecular weights (moles), the following formula will be of benefit:

$$\text{No. of GMWs} = \frac{\text{weight of solute}}{\text{GMW}}$$

(3) Example.

What is the molarity of a solution containing 29.25 grams of sodium chloride in 500 ml. of total solution?

Step 1. Find the number of GMWs.

GMW of NaCl = 58.4 grams

$$\text{No. of GMWs} = \frac{\text{weight of solute}}{\text{GMW}}$$

$$\text{No. of GMWs} = \frac{29.25}{58.4} = 0.5$$

Step 2. Find the molarity.

$$\text{Molarity} = \frac{\text{no. of GMWs of solute}}{\text{no. of liters of solution}}$$

500 ml = 0.5 liter

$$\text{Molarity} = \frac{0.5}{0.5} = \quad \text{1 molar or 1M}$$

7. OXIDATION–REDUCTION REACTIONS (sometimes called redox reaction)

a. Review of Valence.

- Before these reactions are studied, valence should be reviewed briefly.
- The following two valence concepts are especially important in oxidation–reduction reactions:

(1) All elements in their free and uncombined state are considered to have a valence of zero. This holds even for those elements that are diatomic molecules in their free state.

(2) All atoms can exist in a number of valence states. The common valences which you learned previously are the preferred and most stable valences under normal conditions, but other valences can and do occur.

(3) These two concepts are important because oxidation–reduction reactions always involve a change in the valence numbers of some of the elements involved in the reaction.

b. Oxidation.

- Oxidation, in inorganic chemistry, is defined as the loss of electrons or an increase in the valence of an element.
- Consider, for example, the oxidation of elemental iron:

$$Fe^{O}-2e^{-} \dashrightarrow Fe^{+2}$$

- Iron in its free state has a valence of zero and is very reactive since its common valence state is +2 or +3.
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- It loses two electrons to become the ferrous ion.
- The valence has gone from 0 to +2, thus iron has been oxidized. It can undergo further oxidation to the +3 valence state:

$$Fe^{+2} \ -le^{-} \ \text{--->} \ Fe^{+3}$$

- Here the ferrous ion has lost another electron to become a ferric ion.

c. Reduction.

- In inorganic chemistry, reduction is defined as <u>the gain of electrons or a decrease in the valence of an element</u>.
- Consider the reduction of elemental oxygen:

$$O_2 + 4e^- \text{ -----> } 2\,O^{-2}$$

- Observe that oxygen is a diatomic molecule in its free elemental form and has a valence of zero.

$$O_2 + 4e^- \text{ -----> } 2\,O^{-2}$$

- Since the most common valence state of oxygen is –2, oxygen accepts electrons readily to become the oxygen anion.
- The valence of each oxygen atom has gone from 0 to –2, thus oxygen had been reduced.
- If the valence is made smaller (reduced), reduction has occurred.

d. Oxidizing and Reducing Agents.

- For all practical purposes, it is impossible to simply add or subtract electrons from an element except in an electrolytic cell.
- In fact, the oxidation of one element and the reduction of another always occur simultaneously.

- One element loses the electrons; the other element gains the electrons that are lost by the first.
- Consider these two reactions when they are combined:

$$2Fe - 4e^- \text{----->} 2Fe^{+2}$$

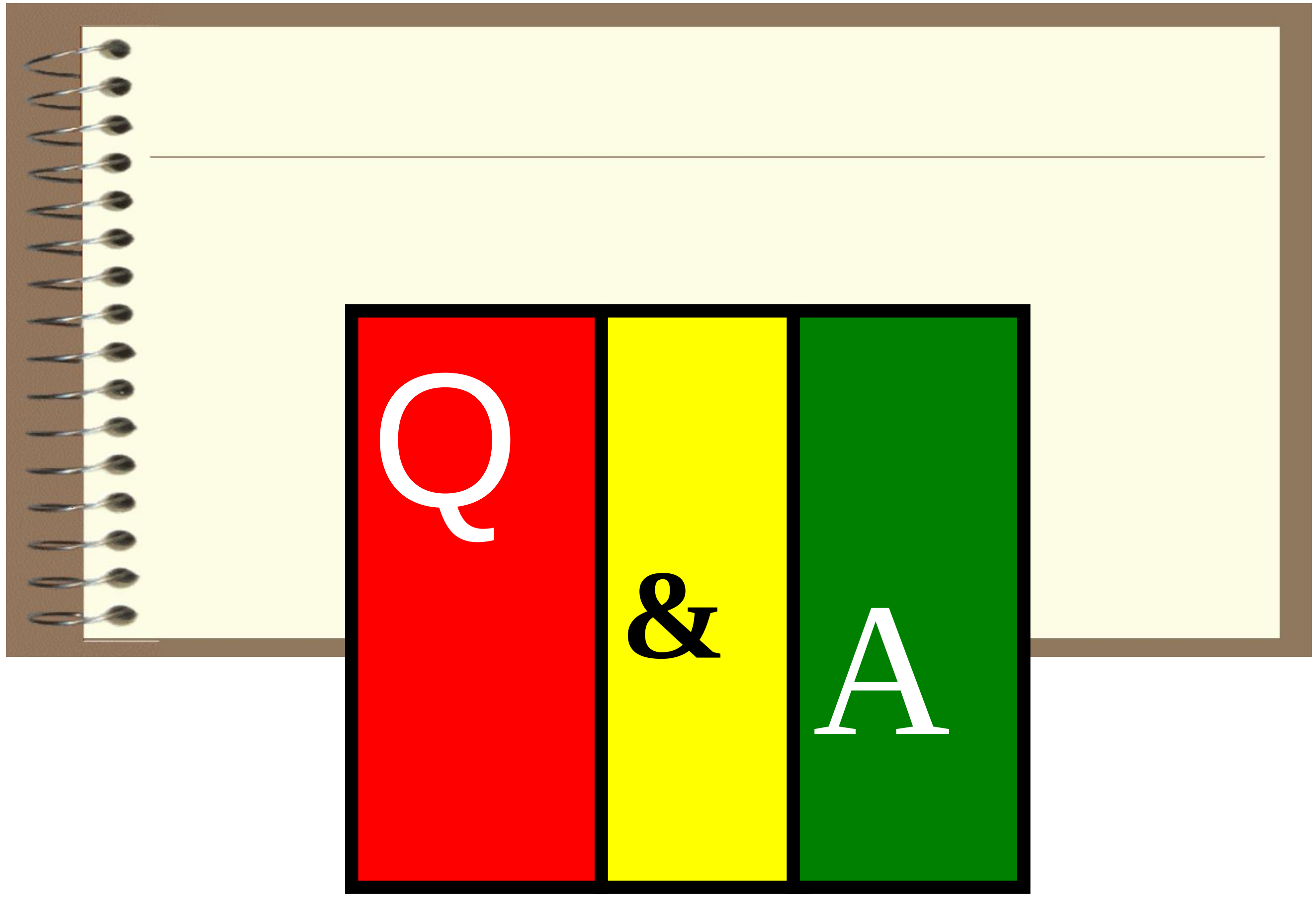
$$O_2 + 4e^- \text{------>} 2O^{-2}$$

$$2Fe + O_2 \text{------>} 2FeO$$

- This is an oxidation–reduction reaction that is very common in our industrialized society.
- The oxidation of iron by atmospheric oxygen gives us iron oxide, commonly known as rust.
- In this reaction, oxygen was reduced, going from a zero to a –2 state by receiving electrons from iron.
- Because it accepted the electrons from iron and allowed the iron to oxidize, oxygen is called an oxidizing agent.

- Iron, which gave up electrons, is called the reducing agent.
- General characteristics of reducing and oxidizing are shown in the following table.

REDUCING AGENT	OXIDIZING AGENT
(1) Gives up electrons	(1) Gains electrons
(2) Oxidized during reaction	(2) Reduced during reaction
(3) Unusually low valence state compared to most common state	(3) Unusually high valence state compared to most common state



Thanks.