Chemistry 104 Chapter Six PowerPoint Notes

Chemical Calculations: Formula Masses, Moles and Chemical Equations
Chapter 6

Chemistry 104
Professor Michael Russell

Chemical Equations

Chemical equation: An expression in which symbols and formulas are used to represent a chemical reaction.

Reactant: A substance that undergoes change in a chemical reaction; written on *left side* of the reaction arrow

Product: A substance that is formed in a chemical reaction; written on *right side* of reaction arrow

$$\underbrace{2\,NaHCO_3}_{\text{Reactant}} \xrightarrow{\text{Heat}} \underbrace{Na_2CO_3 + H_2O + CO_2}_{\text{Products}}$$

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The ***Law of conservation of mass*** states that matter cannot be created or destroyed in any chemical reaction

The bonds between atoms in the reactants are rearranged to form new compounds, but none of the atoms disappear, and no new atoms are formed.

So: Chemical equations must be *balanced*, meaning the numbers and kinds of atoms must be the same on both sides of the reaction arrow.

The numbers placed in front of formulas to balance equations are called *coefficients*, and they multiply all the atoms in the chemical formula.

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Balancing Chemical Equations

The following four steps can be used as a guide to balance chemical equations.

Example: Sulfuric acid reacts with sodium hydroxide to create sodium sulfate and water. Balance this chemical reaction.

Step 1: Write an unbalanced equation, using correct formulas for all reactants and products.

$$H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O \quad \text{(Unbalanced)}$$

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Step 2: Add appropriate coefficients to balance the numbers of atoms of each element.

H_2SO_4 + 2 NaOH ⟶ Na_2SO_4 + H_2O (Balanced for Na)

Add this coefficient . . . . . . to balance these 2 Na.

H_2SO_4 + 2 NaOH ⟶ Na_2SO_4 + H_2O (Balanced for Na and sulfate)

One sulfate here . . . . . . and one here.

Step 3: Check the equation to make sure the numbers and kinds of atoms on both sides of the equation are same.

H_2SO_4 + 2 NaOH ⟶ Na_2SO_4 + 2 H_2O (Completely balanced)

4 H and 2 O here. 4 H and 2 O here.

Step 4: Make sure the coefficients are reduced to their lowest whole-number value (ok here).

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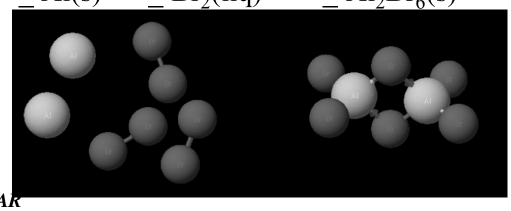
Balancing Equations

_ Al(s) + _ Br_2(liq) ---> _ Al_2Br_6(s)



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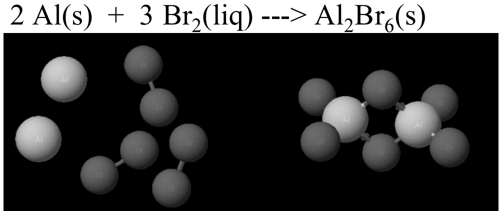
Balancing Equations

2 Al(s) + 3 Br_2(liq) ---> Al_2Br_6(s)



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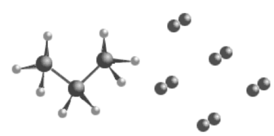

Balancing Equations

____C_3H_8(g) + _____ O_2(g) ---->

_____CO_2(g) + _____ H_2O(g)

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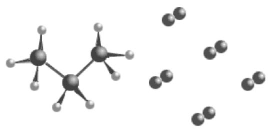

Balancing Equations

$C_3H_8(g)$ + 5 $O_2(g)$ ---->

3 $CO_2(g)$ + 4 $H_2O(g)$

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Balancing Equations - Hints

Balance those atoms which occur in only one compound on each side

Balance the remaining atoms

Reduce coefficients to smallest whole integers

Check your answer

Remember the seven diatomics!

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Test Yourself

Balance the following reactions:

$K_{(s)} + H_2O_{(l)} \rightarrow H_{2(g)} + KOH_{(aq)}$

$Ba_{(s)} + H_3AsO_{4(aq)} \rightarrow H_{2(g)} + Ba_3(AsO_4)_{2(aq)}$

$PCl_{5(s)} + H_2O_{(l)} \rightarrow H_3PO_{4(aq)} + HCl_{(aq)}$

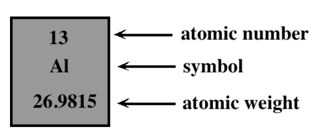
$KClO_{3(s)} \rightarrow KCl_{(s)} + O_{2(g)}$

MAR *practice, practice, practice!*

Atomic Weight - *Ch. 3 Flashback!*

***Atomic Weight*:** The weighted average mass of an element's atoms in a large sample that includes all naturally occurring isotopes of that atom.

Atomic number and atomic weight displayed in periodic table



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Avogadro's Number and the Mole

Molecular weight: The sum of atomic weights for all atoms in a molecule.

- Different molecules have different masses.

Mole: One mole of any substance is the amount whose mass in grams (molar mass) is numerically equal to its molecular or formula weight.

Avogadro's number: The number of molecules in a mole. This number is **6.022×10^{23}.**

We will use the mole to relate our world (grams) to the number of molecules in a substance (Avogadro's number)

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How Big is a Mole?

If you distribute equally 6×10^{23} dollars to every human being on the planet and spend a million dollars an hour it would would take more than 10,000 years to spend your share!

Traveling at the speed of light it would take more than a 100 billion years to travel 6×10^{23} miles!

A stack of 6×10^{23} pennies would be so tall that it would take 100,000 years traveling at the speed of light to go from one end of the stack to the other!

It would take more than a 100 trillion years to print 6×10^{23} dollars at a rate of 100 dollars per second!

Gram – Mole Conversions

Molar mass: The mass in grams of one mole of any substance.

Molar mass = Mass of 1 mole of substance

= Mass of 6.022×10^{23} molecules of substance

= Molecular weight of substance in grams.

Let's calculate the molar mass of water!

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Molar Mass

Example: Find the molar mass of H_2O.

Water has 2 H and 1 O

2*H = 2*1.008 = 2.016 *grams*

1*O = 1*15.999 = 15.999 *grams*

so: Molar mass =

15.999 + 2.016 = **18.015 *grams per mole***

This means that in 18.015 grams of water we have one mole of molecules of water

One mole of water molecules equals 6.022×10^{23} molecules of water

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Molar Mass

What is the molecular weight of Urea, $(NH_2)_2CO$?

Solution:

2 x N = 2 x 14.0067 =	28.0134	
1 x C = 1 x 12.0111 =	12.0111	
4 x H = 4 x 1.00794 =	4.03176	
1 x O = 1 x 15.9994 =	15.9994	
TOTAL =	60.0556 g/mol	

We can convert mol of water to g and g of water to mol using "18.0 g / mol" and dimensional analysis:

Molar mass used as conversion factor

$$0.25 \text{ mol } H_2O \times \frac{18.0 \text{ g } H_2O}{1 \text{ mol } H_2O} = 4.5 \text{ g } H_2O$$

Molar mass used as conversion factor

$$27 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} = 1.5 \text{ mol } H_2O$$

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Test Yourself Part 1

What is the molecular weight of potassium dichromate, $K_2Cr_2O_7$?

Test Yourself Part 2

How many moles in 35.013 g of $K_2Cr_2O_7$?

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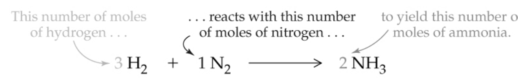
Test Yourself Part 3

How many molecules of $K_2Cr_2O_7$ in 35.013 g?

Coefficients in Chemical Equations

Coefficients in a *balanced* chemical equation tell how many molecules (and thus how many moles) of each reactant are needed *and* how many molecules (and thus moles) of each product are formed.



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Coefficients in Chemical Equations

In the equation:

$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

this equation means

1 mole C_3H_8 + 5 moles O_2 $\rightarrow$ 3 moles CO_2 + 4 moles H_2O

OR

1 molecule C_3H_8 + 5 molecules O_2 $\rightarrow$ 3 molecules CO_2 + 4 molecules H_2O

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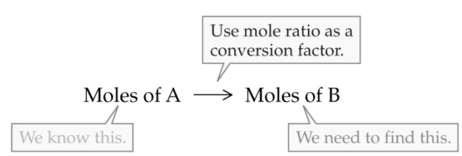
Converting Moles in Equations

We can use a mole ratio from a chemical equation to convert mol (*or g*) of A into mol (*or g*) of B



This is useful in determining how much product is created from so much reactant

Also used for determining how much reactant necessary to create so much product

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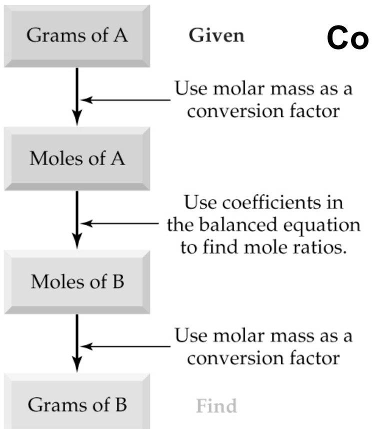
Conversions



Example

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

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Translate Word Equation

Nitrogen + Oxygen → Nitrogen dioxide

⇩ ⇩ ⇩

N_2 O_2 NO_2

Example

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Translate Word Equation

$$N_2 + O_2 \rightarrow NO_2$$

…but note that there are 2 N and 2 O reactants with only 1 N and 2 O products; need to *balance* equation

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Example

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Balance the equation

$$N_2 + \square O_2 \rightarrow \square NO_2$$

Now there are 2 N and 4 O reactants with 2 N and 4 O products; equation is *balanced*

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Convert mass of compound available (nitrogen) to moles.

$$133\ \cancel{g\ N_2} * \frac{1\ mol\ N_2}{28.0\ \cancel{g\ N_2}} = 4.75\ mol\ N_2$$

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Now relate the moles of nitrogen available to moles of oxygen required.

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

$$4.75\ \cancel{mol\ N_2} * \frac{2\ mol\ O_2\ required}{1\ \cancel{mol\ N_2}\ available} =$$

$9.50\ mol\ O_2$ required

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Convert moles of oxygen consumed to grams.

$$9.50\ \cancel{mol\ O_2}\ required * \frac{32.0\ g\ O_2}{1\ \cancel{mol\ O_2}} =$$

$304\ g\ O_2$ required to combust $133\ g\ N_2$

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Relate the moles of nitrogen available to moles of NO_2 produced.

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

$$4.75\ \text{mol}\ N_2 * \frac{2\ \text{mol}\ NO_2\ \text{produced}}{1\ \text{mol}\ N_2\ \text{available}} =$$

9.50 mol NO_2 produced

Example:

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

Convert moles of NO_2 produced to grams.

$$9.50\ \text{mol}\ NO_2 * \frac{46.0\ g\ NO_2}{1\ \text{mol}\ NO_2} =$$

437g NO_2 created when 133 g N_2 burned in O_2

Test Yourself

How many grams of water will be produced if $6.000*10^9$ molecules of sulfur dioxide are created? The equation:

$$2\ H_2SO_4 + Cu \rightarrow SO_2 + 2\ H_2O + CuSO_4$$

Percent Yield

Actual yield - The quantity of product (g) actually obtained in a reaction.

Theoretical yield - The quantity of product (g) that is expected from a chemical reaction.

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Example - Percent Yield

Nitrogen and oxygen gases combine to give nitrogen dioxide. How many grams of oxygen are consumed and how many grams of nitrogen dioxide can be formed by the reaction of 133 g of nitrogen in excess oxygen?

The theoretical yield of NO_2 was 434 g (see previous example)

In the actual experiment, only 247 g of NO_2 was recovered; this is the actual yield

Percent yield = (actual / theoretical) * 100%
= (247 / 434) * 100%
= **56.9%**

Practice, practice, practice!

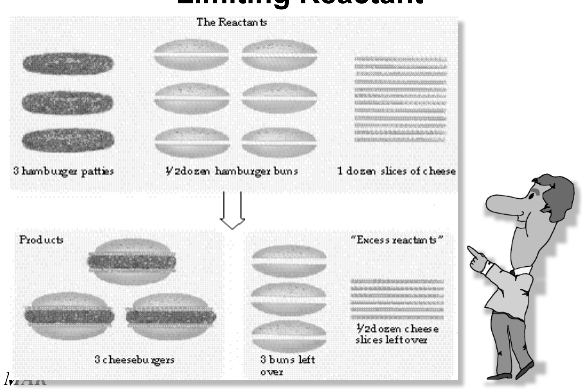
Limiting Reactant

Most of the time, one reactant limits how much product can be produced.

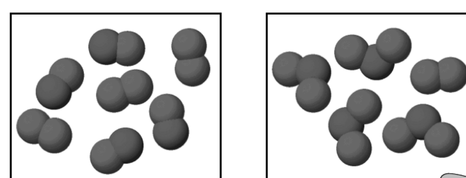
This reactant is called the **limiting reactant**, and the other reactant(s) is called the **reactant in excess** or **excess reactant**.

The limiting reactant ***limits*** the **amount of product** that can be made and hence controls the reaction.

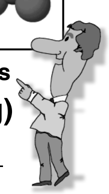
Limiting Reactant



Limiting Reactant



Reactants Products



$$2\ NO(g) + O_2\ (g) \rightarrow 2\ NO_2(g)$$

Limiting reactant = ___________

Excess reactant = ____________

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Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make $ZnCl_2$ and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of $ZnCl_2$.

balanced chemical equation:

$$Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make $ZnCl_2$ and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of $ZnCl_2$. $Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

First, convert Zn to mol to use mol ratio:

7.00 g Zn * (mol Zn / 65.4 g Zn) = 0.107 mol Zn

Second, convert mol Zn and mol HCl to mol $ZnCl_2$.

0.107 mol Zn * (mol $ZnCl_2$ / mol Zn) = 0.107 mol $ZnCl_2$

0.100 mol HCl * (mol $ZnCl_2$ / 2 mol HCl) = 0.0500 mol $ZnCl_2$

Notice quantities of $ZnCl_2$ created!

Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make $ZnCl_2$ and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of $ZnCl_2$. $Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

Third, compare $ZnCl_2$ quantities:

7.00 g Zn = 0.107 mol Zn which gives 0.107 mol $ZnCl_2$

0.100 mol HCl gives 0.0500 mol $ZnCl_2$

HCl gives *less* $ZnCl_2$ than Zn! Hence, HCl is *limiting* how much product can be made, and **HCl is the limiting reactant**.

Zn could produce a lot more $ZnCl_2$ than HCl, but HCl cannot keep up. Hence, **Zn is the reactant in excess**.

Example - Limiting Reactant

Zinc metal (7.00 g) reacts with 0.100 mol HCl to make $ZnCl_2$ and H_2 gas. Find the limiting reactant, the reactant in excess, and find the theoretical yield of $ZnCl_2$. $Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

Fourth, find theoretical yield of $ZnCl_2$.

Need to use limiting reactant to calculate theoretical yield of $ZnCl_2$.

0.0500 mol $ZnCl_2$ * (136.3 g $ZnCl_2$ / mol $ZnCl_2$) = **6.82 g $ZnCl_2$**, the **theoretical yield**

If we used the 0.107 mol $ZnCl_2$ from Zn:

0.107 mol $ZnCl_2$ * (136.3 g $ZnCl_2$ / mol $ZnCl_2$) = 14.6 g $ZnCl_2$, *which is not possible!*

Practice, practice, practice!

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Test Yourself Part 1

Write a balanced reaction for the formation of gaseous antimony(III) chloride from gaseous antimony and chlorine gas.

Test Yourself Part 2

Determine the limiting reactant and theoretical yield of $SbCl_3$ if 129 g of Sb and 106 g of Cl_2 are mixed.

$2\ Sb_{(g)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(g)}$

Test Yourself Part 3

Only 113.5 g of $SbCl_3$ were collected. Calculate the percent yield for the reaction. (Theoretical yield = 227 g)

$2\ Sb_{(g)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(g)}$

Types of Chemical Reactions

Chemical reactions are grouped into three general classes:

Precipitation reactions: Processes in which an insoluble solid (a *precipitate*) forms when reactants are combined in aqueous solution.

Acid-base (neutralization) reactions: Processes in which an *acid* reacts with a *base* to yield *water* plus an ionic compound called a *salt*.

Oxidation-reduction (Redox) reactions: Processes in which one or more *electrons* are transferred between reaction partners.

Several kinds of redox reactions: *decomposition*, *combination* and *single replacement - see lab*

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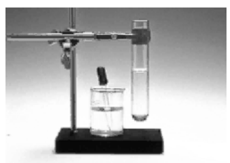
Precipitation Reactions: Solubility Guidelines

Solubility: The amount of a compound that will dissolve in a given amount of solvent at a given temperature.

When solubility exceeded, precipitates form



$Pb^{2+}_{(aq)} + 2\,I^{-}_{(aq)} \rightarrow PbI_{2(s)}$

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Precipitation Reactions: Solubility Guidelines

Useful solubility guidelines for CH 104:

- A compound is *probably soluble* if it contains one of the following *cations*:
  - Group 1A cations: Li^+, Na^+, K^+, Rb^+, Cs^+
  - Ammonium ion, NH_4^+
- A compound is *probably soluble* if it contains one of the following *anions*:
  - Halides: Cl^-, Br^-, I^- *except* Ag^+, Hg_2^{2+}, and Pb^{2+} compounds.
  - Nitrate (NO_3^-), perchlorate (ClO_4^-), acetate ($CH_3CO_2^-$)
  - sulfate (SO_4^{-2}) *except* Ba^{2+}, Hg_2^{2+}, and Pb^{2+} sulfates.

Examples: NaCl, K_2SO_4 - *soluble*

Examples: AgCl, $BaSO_4$ - *insoluble*

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Acids, Bases & Neutralizations

When equal amounts (moles) of acids (H^+) and bases (OH^-) are mixed together, both acidic and basic properties disappear because of a neutralization reaction. The neutralization reaction produces water and a salt.

Example: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

acid *base* *water* *salt*

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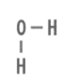
Acids, Bases & Neutralizations

Example: $HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

acid *base* *water* *salt*



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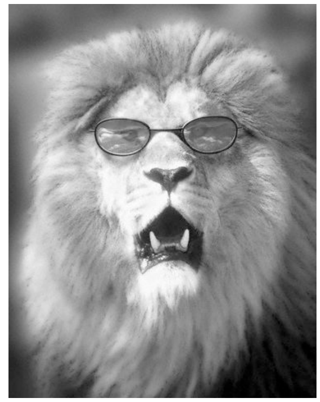
Oxidation-Reduction

Oxidation-Reduction reactions, often called 'Redox Reactions', are another way to classify chemical reactions.

Redox reactions are simply classified as reactions where the oxidation number or "charge" of the species involved changes during the chemical reaction.

Redox reactions are characterized by *Electron Transfer* between an electron donor and electron acceptor.



LEO says GER

LEO says GER

Lose	**G**ain
Electrons	**E**lectrons
Oxidized	**R**educed

$Zn(s) \rightarrow Zn^{2+} + 2e^-$ ***Oxidized***

$Cu^{2+} + 2e^- \rightarrow Cu(s)$ ***Reduced***

Redox Reactions

Oxidation: The loss of one or more electrons by an atom (LEO)

Reduction: The gain of one or more electrons by an atom (GER)

Redox reaction: A reaction in which one or more electrons are transferred from one atom to another.

Reducing agent: The species undergoing oxidation; causes reduction

Oxidizing agent: The species undergoing reduction; causes oxidation

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Recognizing Redox Reactions

One can determine if a reaction is redox by comparing oxidation numbers

An ***Oxidation number*** indicates whether an atom is neutral, electron rich, or electron poor.

By comparing the oxidation number of an atom before and after reaction, we can tell whether the atom has gained or lost electrons.

Non-redox reactions (precipitation, acid-base) have no change in oxidation number

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Rules for Oxidation Numbers

Atoms in elemental states have oxidation number of 0.

Example: Na, Br_2, H_2 all have 0 oxidation number

Monoatomic ions have oxidation number equal to its charge.

Example: Na^+: ox. # = +1

Example: S^{2-}: ox. # = -2.

In molecular compounds, atoms usually have the same oxidation number if it were a monatomic ion. The sum of oxidation numbers in a compound is equal to its ionic charge.

Example: in H_2O, each H has oxidation number +1 and oxygen has oxidation number of -2; sum = 0

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Rules for Oxidation Numbers

Example: Find oxidation numbers for each element in NH_4^+.

Hydrogen ox. # = +1

Nitrogen ox. # = -3

Sum of ox. #s = -3 + 4*(+1) = +1, which equals charge on NH_4^+ (+1), ok!

Example: Find oxidation numbers for each element in Na_2S.

Sodium ox. # = +1

Sulfur ox. # = -2

Sum of ox. #s = -2 + 2*(+1) = 0, which equals charge on Na_2S (neutral), ok!

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Recognizing Redox - Oxidation

Oxidation is the loss of electrons by a substance undergoing a chemical reaction or, in other words, the change in oxidation number of a substance to a more positive oxidation number.

$$4\ Fe^0 + 3\ O_2 \rightarrow 2\ Fe_2^{3+}O_3$$

oxidizing agent

Fe → Fe^{3+} + 3 e$^-$ *oxidation*

Recognizing Redox - Reduction

Reduction is the gain of electrons by a substance undergoing a chemical reaction or, in other words, the change in oxidation number of a substance to a more negative oxidation number.

$$4\,Fe + 3\,O_2^{0} \rightarrow 2\,Fe_2O_3^{2-}$$

reducing agent

$$^{1}/_{2}\,O_2 + 2\,e^- \rightarrow O^{2-}$$ *reduction*

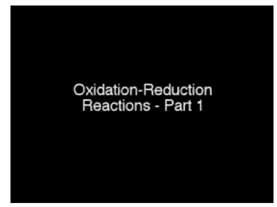
This type of redox reaction called a **combination** reaction

Example: Redox Reaction

Zinc metal and oxygen gas combine to make zinc(II) oxide. Determine which substance gets oxidized and which substance gets reduced.

Balanced reaction:

$$2\,Zn + O_2 \rightarrow 2\,ZnO$$



Example: Redox Reaction

Magnesium metal and oxygen gas combine to make magnesium oxide. Determine which substance gets oxidized and which substance gets reduced.

Balanced reaction:

$$2\,Mg + O_2 \rightarrow 2\,MgO$$

Goal: determine if Mg and O_2 are oxidized or reduced

Example: Redox Reaction

Use your knowledge of the periodic table to determine the oxidation states of each of the species.

- **Remember that species in their 'natural' state (uncombined) have oxidation numbers of zero.**

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Example: Redox Reaction

Use your knowledge of the periodic table to determine the oxidation states of each of the species.

Remember that species in their 'natural' state (uncombined) have oxidation numbers of zero.

Therefore:

$$2\ Mg^{0} + O_2 \rightarrow 2\ Mg^{2+}O$$

Example: Redox Reaction

Use your knowledge of the periodic table to determine the oxidation states of each of the species.

Remember that species in their 'natural' state (uncombined) have oxidation numbers of zero.

Therefore:

$$2\ Mg^{0} + O_2^{0} \rightarrow 2\ Mg^{2+}O^{2-}$$

Example: Redox Reaction

$$2\ Mg^{0} + O_2 \rightarrow 2\ Mg^{2+}O$$

The oxidation number of Mg changes from
$0 \rightarrow 2+$

Thus, **Mg** is **Oxidized,** and Mg is the **reducing agent.**

Example: Redox Reaction

$$2\ Mg + O_2^{0} \rightarrow 2\ MgO^{2-}$$

The oxidation number of O_2 changes from
$0 \rightarrow 2-$

Thus, **O_2** is **Reduced,** and O_2 is the **oxidizing agent.**

This type of redox reaction is called a **combination** reaction

Practice, practice, practice!

Chemistry 104 Chapter Six PowerPoint Notes

Other Redox Reactions

We have already seen **combination** redox reactions:

$$2\ Na + Cl_2 \rightarrow 2\ NaCl$$

Another type of redox reaction: **decomposition**

$$2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$$

Another type of redox reaction: **single replacement**

$$CuCl_2 + Zn \rightarrow ZnCl_2 + Cu$$

Watch lab and problems for more examples!

End of Chapter 6

To review and study for Chapter 6, look at the "Concepts to Remember" at the end of Chapter Six