Brief Review for 1311 Honors Exam 2

- Chapter 2: Periodic Table
  1. Metals
     1. Representative Metals
        Alkali Metals ………………… Group 1
        Alkaline Earth Metals …….. Group 2
     2. Transition Metals
  II. Metalloids
  II. NonMetals

- Chapter 3: All
- Chapter 4: All

Metals, Semimetals (Metalloids) & Nonmetals

1. Which of the following is a metal
   (a) S (b)Si (c) Sr (d) Se (e) P
2. Which of the following is a nonmetal
   (a) As (b)Ti (c) B (d) Se (e) Pb
3. All of the following are metalloids except
   (a) B (b) As (c) Al (d) Ge (e) Si

Chapter 3: Chemical Stoichiometry

- Chemical Equations – (Write, Balance, Interpret)
- Reactions You Should Know
- Formula Weights – (Must know chemical formula)
- Avogadro’s number and the Mole
- Limiting Reactants
- Per Cent Yield
- Empirical Formulas
- Solubility Rules
- Oxidation Rules

Chemical REACTIONS you should know

Combustion (Of a Hydrocarbon)
C₂H₅OH + O₂ \rightarrow CO₂ + H₂O

Neutralization (Acid + Base)
HC₂H₃O₂(aq) + NaOH(aq) \rightarrow H₂O + NaC₂H₃O₂(aq)

Precipitation
BaCl₂(aq) + Na₂SO₄(aq) \rightarrow BaSO₄(s) + NaCl(aq)

Chemical Reactions

Combustion of Ethanol
1st Write Reaction
C₂H₅OH + O₂ \rightarrow CO₂ + H₂O

2nd Balance Equation
1 C₂H₅OH + ? O₂ \rightarrow 2 CO₂ + 7/2 H₂O

must start somewhere

why not here

then and

What number for oxygen?  

A balanced chemical equation has the same number of **atoms** in the reactants as in the products.

\[ 4 \text{C}_2\text{H}_6\text{OH} + 13 \text{O}_2 \rightarrow 8 \text{CO}_2 + 14 \text{H}_2\text{O} \]

**3rd Interpret Reaction**

The reactants must be in the ratio of \(4 : 13\)

The balanced equation has

\(4 + 13 + 8 + 14 = 39\) total moles

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### Calculate Formula Weights

**Must Know Formulas**

- **Oxygen** \(\text{O}_2 : 2 \times 16 = 32\) grams per unit
- **Ozone** \(\text{O}_3 : 3 \times 16 = 48\) grams per mole
- **Methane** \(\text{CH}_4 : 12 + 4 \times 1 = 16\) gms / mole
- **Ammonia** \(\text{NH}_3 : 14 + 3 \times 1 = 17\) g / mole
- **C\(_6\)H\(_12\)O\(_6\)** \((6 \times 12) + (12 \times 1) + (6 \times 16) = 180\)

### Use Formula Weight to Convert

**Grams to Moles & Moles to Grams**

- **How many moles in 2.4 grams of Ozone**
  
  \[2.4 \text{ Grams} \times \frac{1 \text{ moles}}{48 \text{ Grams}} = \text{ moles}\]

  Must know formula for Ozone!

  \[2.4 \text{ Grams} \times \frac{1 \text{ moles}}{48 \text{ Grams}} = 0.050 \text{ moles}\]

### Use AVOGADRO’S Number to Convert

**Molecules to Moles & Moles to Molecules**

- **How many Hydrogen atoms in 1 mole of Methane**

  \[ \frac{6.02 \times 10^{23}}{} \times \frac{4 \text{ hydrogen atoms}}{1 \text{ molecule}} = 2.41 \times 10^{24} \text{ hydrogen atoms} \]

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### Use Chemical Formula to Convert

**Molecules to Atoms & Atoms to Molecules**

- **How many Hydrogen atoms in 1 mole of Methane**

  \[ \frac{6.02 \times 10^{23} \text{ molecules}}{} \times \frac{4 \text{ hydrogen atoms}}{1 \text{ molecule}} = 2.41 \times 10^{24} \text{ hydrogen atoms} \]
Interpretation of Chemical Reactions Using Stoichiometry

- The chemical arithmetic necessary to relate the moles of reactants & products
- Chemical Reactions are Interpreted on the Mole basis but chemicals are weighed in the laboratory in Grams

Yields of Chemical Reactions

- Chemical Reactions do not always go the way we expect them to
- Using stoichiometry we can calculate the theoretical (Maximum) amount of product formed in a reaction.

If 10.0 moles of NO are reacted with 5.0 mole O₂, how many moles NO₂ are produced?

\[ 2 \text{NO (g)} + \text{O}_2 (g) \rightarrow 2 \text{NO}_2 (g) \]

(a) 2.0 mol NO₂
(b) 6.0 mol NO₂
(c) 10.0 mol NO₂
(d) 16.0 mol NO₂
(e) 32.0 mol NO₂

The Limiting Reactant

A reaction stops when one reactant is totally consumed.

This is the limiting reactant.
The other reactants are excess reactants.

If 10.0 moles of NO are reacted with 6.0 moles O₂, how many moles of the excess reagent remain?

- 1.0 mol O₂
- 5.0 mol O₂
- 4.0 mol NO
- 8.0 mol NO

If 10.0 moles of NO are reacted with 6.0 moles O₂, how many moles NO₂ are produced?

\[ 2 \text{NO (g)} + \text{O}_2 (g) \rightarrow 2 \text{NO}_2 (g) \]

(a) What LIMITS the reaction? The amount of NO
(b) What is in excess? The amount of Oxygen
Yields of Chemical Reactions

- Using stoichiometry the **theoretical** (Maximum) amount of product formed is calculated.
- If the actual amount of product formed is less than the theoretical amount, the **percentage yield** is calculated.

\[
\text{% yield} = \frac{\text{Actual product yield}}{\text{Theoretical product yield}} \times 100\%
\]

With a 50% Yield, How many moles of NH₃ are produced from 1 ½ moles of H₂ and 1 mole of N₂?

1st Write and Balance Reaction:

\[
3 \text{ H}_2 + 1 \text{ N}_2 \rightarrow 2 \text{ NH}_3
\]

or

\[
1 \frac{1}{2} \text{ H}_2 + \frac{1}{2} \text{ N}_2 \rightarrow 1 \text{ NH}_3
\]

So ½ mole NH₃ formed with ½ mole of Nitrogen in excess

If 0.720 grams O₃ reacts with 0.600 grams of nitrogen monoxide, to produce nitrogen dioxide

1. How many grams of product will be formed?
2. What is the limiting reagent?
3. How much excess reagent will remain?

1st Write & Balance Reaction:

\[
1 \text{ O}_3 + 3 \text{ NO} \rightarrow 3 \text{ NO}_2
\]

6 = O = 6
3 = N = 3

0.720 grams O₃ & 0.600 grams NO

For O₃: 0.720 grams × \(\frac{1 \text{ mole}}{48 \text{ grams}}\) = 0.015 moles O₃

For NO have 0.600 grams × \(\frac{1 \text{ mole}}{30.0 \text{ grams}}\) = 0.020 moles

\[
1 \text{ O}_3 + 3 \text{ NO} \rightarrow 3 \text{ NO}_2
\]

1 (0.015) O₃ + 3 (0.015) NO → 3 (0.015) NO₂

0.015 O₃ + 0.045 NO → 0.045 NO₂

0.015 \text{ O}_3 + 0.045 \text{ NO} \rightarrow 0.045 \text{ NO}_2

Therefore Nitrogen monoxide is limiting reagent

Two Approaches to Chemical Formulas

1. **Given Formula** Determine % Composition of Each Element in Compound.

   *What is the % H & O in hydrogen peroxide?*

2. **Determine Formula** Given % Composition of Each Element in Compound

   *What is the Formula of the compound that is 5.88% H & 94.12% O?*
Find Empirical Formula

Weight K = 0.421
Weight O = 0.084
Total Weight = 0.505
First determine % by Weight

\[
\% \text{ by Wt K} = \frac{0.421}{0.505} \times 100 = 83.36633
\]

\[83.4\% \text{ by wt K and } 16.6\% \text{ by wt O}\]

Choose any total weight 100 grams is convenient
Construct the following work sheet

<table>
<thead>
<tr>
<th>Element</th>
<th>%</th>
<th>Wt</th>
<th>moles</th>
<th>ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>K</td>
<td>83.4</td>
<td>83.4 g</td>
<td>(\frac{83.4 \text{ g}}{39}) = 2.14 = 2</td>
<td></td>
</tr>
<tr>
<td>O</td>
<td>16.6</td>
<td>16.6 g</td>
<td>(\frac{16.6 \text{ g}}{16}) = 1.04 = 1</td>
<td></td>
</tr>
</tbody>
</table>

Therefore formula is \(\text{K}_2\text{O}\)

Chapter 4
Reactions in Aqueous Solutions & Solution Stoichiometry

WHAT IS A SOLUTION?

A HOMOGENEOUS MIXTURE OF TWO OR MORE SUBSTANCES
Solvent + Solute = Solution

Concentrations in Solutions Molarity

ELECTROLYTES

- **ELECTROLYTE** – A substance that dissolves in water to produce IONS
Example: HCl(aq), NaOH(aq), NaCl(aq)

- **NON ELECTROLYTE** A substance that DOES NOT produce IONS, remain as molecules, when dissolved in water.
Example: sugar(aq), Ethylene glycol (aq)

<table>
<thead>
<tr>
<th>Table 4.1 Solubility Guidelines for Common Ionic Compounds in Water</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Soluble Ionic Compounds</strong></td>
</tr>
<tr>
<td>Compounds containing (\text{NO}_3^-)</td>
</tr>
<tr>
<td>(\text{C}_2\text{H}_5\text{O}_2^-)</td>
</tr>
<tr>
<td>(\text{Cl}^-)</td>
</tr>
<tr>
<td>(\text{Br}^-)</td>
</tr>
<tr>
<td>(\text{I}^-)</td>
</tr>
<tr>
<td>(\text{SO}_4^{2-})</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th><strong>Insoluble Ionic Compounds</strong></th>
<th><strong>Important Exceptions</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing (\text{S}^{2-})</td>
<td>Compounds of (\text{NH}_4^+), the alkali metal cations, and (\text{Ca}^{2+}), (\text{Sr}^{2+}), and (\text{Ba}^{2+})</td>
</tr>
<tr>
<td>(\text{CO}_3^{2-})</td>
<td>Compounds of (\text{NH}_4^+) and the alkali metal cations</td>
</tr>
<tr>
<td>(\text{PO}_4^{3-})</td>
<td>Compounds of (\text{NH}_4^+) and the alkali metal cations</td>
</tr>
<tr>
<td>(\text{OH}^-)</td>
<td>Compounds of the alkali metal cations, and (\text{Ca}^{2+}), (\text{Sr}^{2+}), and (\text{Ba}^{2+})</td>
</tr>
</tbody>
</table>
What do the following compounds do when mixed with water? (are they water soluble?)

<table>
<thead>
<tr>
<th>Compound</th>
<th>Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium Iodide</td>
<td>( \text{MgI}_2 \rightarrow \text{Mg}^{2+}(aq) + 2\text{I}^-(aq) )</td>
</tr>
<tr>
<td>Aluminum Nitrate</td>
<td>( \text{Al(NO}_3)_3 \rightarrow \text{Al}^{3+}(aq) + 3\text{NO}_3^-(aq) )</td>
</tr>
<tr>
<td>Ammonium Sulfate</td>
<td>( (\text{NH}_4)_2\text{SO}_4 \rightarrow 2\text{NH}_4^+(aq) + \text{SO}_4^{2-}(aq) )</td>
</tr>
<tr>
<td>Perchloric Acid</td>
<td>( \text{HClO}_4 \rightarrow \text{H}^+(aq) + \text{ClO}_4^-(aq) )</td>
</tr>
<tr>
<td>Silver Chloride</td>
<td>AgCl Is Insoluble</td>
</tr>
</tbody>
</table>

WHY DOES A REACTION OCCUR?

1. Precipitation Reaction
   - An INSOLUBLE Solid is Formed
2. Neutralization Reaction
   - Acid-Base slightly ionized Substance water Formed
3. Oxidation-Reduction Reaction
   - Change in Oxidation State

Will precipitation occur when the following solutions are mixed?

- If so, write net ionic equation for reaction
- (a) \( \text{Na}_2\text{CO}_3 \) and \( \text{Ag C}_2\text{H}_3\text{O}_2 \) ……… Yes
  \( 2\text{Ag}^{+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{Ag}_2\text{CO}_3(s) \)
- (b) \( \text{NaNO}_3 \) and \( \text{NiSO}_4 \) ………… No reaction
- (c) \( \text{FeSO}_4 \) and \( \text{Pb(ClO}_4)_2 \) …………. Yes
  \( \text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) \)

CHEMICAL EQUATIONS

- **MOLECULAR** - Uses The Full Formula Of Reactants and Products
- **IONIC** – ONLY Ions Are Shown
- **NET IONIC** – SPECTATOR Ions Are Removed From Ionic Equation
- **SPECTATOR Ions** – Ions that do not take part in the chemical reaction

Equations for Silver Perchlorate + Sodium Nitrate

**The Molecular Equation is:**
\( \text{AgClO}_4(aq) + \text{NaNO}_3(aq) \rightarrow \text{AgNO}_3(aq) + \text{NaClO}_4(aq) \)

**The Ionic Equation is:**
\( \text{Ag}^+(aq) + \text{ClO}_4^-(aq) + \text{Na}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{Ag}^+(aq) + \text{NO}_3^-(aq) + \text{Na}^+(aq) + \text{ClO}_4^-(aq) \)

**Spectator Ions:** all are spectator ions

**There is no Net Ionic Equation because there is NO REACTION !!!!!**

Write (1) Molecular (2) Ionic & (3) Net Ionic Equations for

\( \text{H}_2\text{SO}_4(aq) + \text{KOH}(aq) \rightarrow \) ????
\( \rightarrow \text{H}_2\text{O}(liq) + \text{K}_2\text{SO}_4(aq) \)
\( 2\text{H}^+(aq) + \text{SO}_4^{2-}(aq) + \text{K}^+(aq) + \text{OH}^-(aq) \rightarrow \) 
\( \rightarrow \text{H}_2\text{O}(liq) + \text{K}_2\text{SO}_4(aq) \)
\( 2\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(liq) \)
Equations for Sulfuric Acid + Potassium Hydroxide

The Molecular Equation is:
\[ \text{H}_2\text{SO}_4(\text{aq}) + 2 \text{KOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{liq}) + \text{K}_2\text{SO}_4(\text{aq}) \]

The Ionic Equation is:
\[ 2 \text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2 \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \]
\[ \text{H}_2\text{O}(\text{liq}) + 2 \text{K}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \]

Spectator Ions K⁺ & SO₄²⁻

Net Ionic Equation is:
\[ 2 \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{liq}) \]

How do you recognize an Oxidation Reduction Reaction?

H₂ (g) + O₂ (g) \rightarrow H₂O (g)

On “one side” of \rightarrow element all by itself
While on the “other side” element in combination

Another example
Zn (s) + 2 HCl (aq) \rightarrow H₂ (g) + ZnCl₂ (aq)

Assigning Oxidation Numbers

Must Know Oxidation “RULES” See Text / Lab Manual

1. EACH ATOM in a PURE ELEMENT has an Oxidation Number of ZERO
   \text{H}_2, \text{O}_2, \text{N}_2, \text{F}_2, \text{Na}, \text{Fe}, \text{Zn} etc

   For example
   Sodium .................Na
   Atomic Number ........11
   Number of Protons ... 11
   Number of Electrons ...11
   Same number of protons and electrons

2. Alkali Metals always = +1 {Li, Na, K, Rb, Cs}
3. Alkaline Earth Metals always = +2 {Mg, Ca, Sr, Ba}
4. The oxidation number of H in its compounds = +1
   Most of the time: \text{H}_2\text{O}; \text{NaOH}; \text{HCl}; \text{H}_2\text{C}_2\text{O}_4
   \text{But Not all the time} Ca H₂ {H = -1 in Hydrides}
5. Oxidation number of F in compounds always = -1
6. Oxidation number of O in its compounds = -2
   Most of the time: \text{H}_2\text{O}; \text{NaOH}; \text{H}_2\text{C}_2\text{O}_4
   \text{But Not all the time} H₂O₂ {O = -1 in Peroxides
OXIDATION REDUCTION

• What is The OXIDATION NUMBER For EACH Element Present?
• What is OXIDIZED?
• What is REDUCED?
• What is The OXIDIZING AGENT?
• What is The REDUCING AGENT?

“Types” of Equations

(a) Molecular: the formula for each substance are written in the molecular form

\[ \text{Zn (s) + 2 HBr (aq)} \rightarrow \text{H}_2 (g) + \text{ZnBr}_2 (aq) \]

(b) Ionic: shows dissolved species in ionic form

\[ \text{Zn(s) + 2H}^+(aq) + 2\text{Br}^-(aq) \rightarrow \text{H}_2(g) + \text{Zn}^{2+}(aq) + 2\text{Br}^-(aq) \]

(c) Net Ionic: only shows the ionic species that actually take part in the reaction

\[ \text{Zn (s) + 2 H}^+(aq) \rightarrow \text{H}_2(g) + \text{Zn}^{2+}(aq) \]

Redox Reactions

\[ \text{Zn(s) + HCl(aq)} \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq) \]

Oxidation Numbers

\[ \begin{align*}
\text{Zn(s) & H}_2(g) &= 0 \\
\text{Zn in ZnCl}_2(aq) &= +2 \\
\text{Cl in HCl and ZnCl}_2 &= -1
\end{align*} \]

What is OXIDIZED

Zn \( ^0 \) \( \rightarrow \) Zn \( ^{2+} \)

What is REDUCED

H \( ^+ \) \( \rightarrow \) H\( _2 \) \( ^0 \)

Oxidizing AGENT?

HCl(aq)

Reducing AGENT?

Zn metal

Putting Concepts Together

A sample of 70.5 mg of potassium phosphate is added to 15.0 mL of 0.050 \( M \) silver nitrate,

(a) Write the molecular equation for the reaction.

(b) Write ionic and net ionic equations

(c) What is the limiting reactant in the reaction?

(d) Calculate the theoretical yield, in grams, of the precipitate that forms.

Molecular Equation

K\(_3\)PO\(_4\)(aq) + AgNO\(_3\)(aq) \( \rightarrow \) Ag\(_3\)PO\(_4\)? + KNO\(_3\)?

According to solubility guidelines Ag\(_3\)PO\(_4\) will precipitate while K\(_3\)PO\(_4\), AgNO\(_3\), & KNO\(_3\) are soluble

Complete ionic equation

K\(^+\)(aq) + PO\(_4^{3-}\)(aq) + Ag\(^+\)(aq) + NO\(_3^-\)(aq) \( \rightarrow \) Ag\(_3\)PO\(_4\)(s) + K\(^+\)(aq) + NO\(_3^-\)(aq)

ionic equation

K\(^+\)(aq) + PO\(_4^{3-}\)(aq) + Ag\(^+\)(aq) + NO\(_3^-\)(aq) \( \rightarrow \) Ag\(_3\)PO\(_4\)(s)

Net Ionic Equation

Ag\(^+\)(aq) + PO\(_4^{3-}\)(aq) \( \rightarrow \) Ag\(_3\)PO\(_4\)(s)

Spectator ions: K\(^+\) & NO\(_3^-\) ions

determine the number of moles of each reactant.

K\(^-\)(aq) + Ag\(^+\)(aq) + NO\(_3^-\)(aq) \( \rightarrow \) Ag\(_3\)PO\(_4\)(s) + K\(^+\)(aq) + NO\(_3^-\)(aq)
According to the balanced equation, 
\[ \text{K}_3\text{PO}_4(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}_3\text{PO}_4 + 3\text{KNO}_3(\text{aq}) \]
1 mol \( \text{K}_3\text{PO}_4 \) requires 3 mol of \( \text{AgNO}_3 \).

There are \( 3.32 \times 10^{-4} \) moles of \( \text{K}_3\text{PO}_4 \) and \( 7.5 \times 10^{-4} \) moles of \( \text{AgNO}_3 \).

Rebalance equation:
\[ 3.32 \text{K}_3\text{PO}_4 + 9.96\text{AgNO}_3 \rightarrow 3.32\text{Ag}_3\text{PO}_4 + 9.96\text{KNO}_3 \]
Not enough \( \text{AgNO}_3 \) \( (\text{have } 7.5 \times 10^{-4} \text{ moles}) \) for the \( \text{K}_3\text{PO}_4 \), therefore \( \text{AgNO}_3 \) is the limiting reactant.

**Molarity (M)**

**Memorize the definition**

\[ \text{Molarity} = \frac{\text{moles of solute}}{\text{Liters of solution}} \]

Moles Solute = Molarity \( \times \) Liters of solution

\[ \text{moles} = \text{M} \times V \]

**How are Moles determined from Molarity?**

Moles of Solute = Molarity \( \times \) (Volume in Liters)

Calculate the number of moles of HCl in 50.0 mL of 2.00 M HCl(aq)
Moles = M \( \times \) V = \( (0.0500) \times (2.00) = 0.100 \)

Calculate the number of moles of NaOH in 51.0 mL of 2.00 M NaOH
Moles = M \( \times \) V = \( (0.0510) \times (2.00) = 0.102 \)

**Preparing solutions By Dilution**

Start with 50.0 mL of 2.00 M HCl(aq)
Add water until have 100.0 mL of solution
What is the Molarity of the new solution?

1. \( M_1 \times V_1 = M_2 \times V_2 \) = moles of Solute
2. \( (2.00)(50.0) = (M_2)(100.0) = ??? \)
3. \( M_2 = 1.00 \)

**Chap 5 THERMO chemistry Review**

**HEAT CHANGES** that take place during

1. Physical Changes (Change of Phase) &
2. Chemical Changes (Chemical Reactions)

**Energy cannot be created or destroyed**

BUT it can be transformed from one form to another
(potential & kinetic)
or transferred between system and surroundings

One of the basic assumptions of thermodynamics is the idea that the Universe can be divided into a System and its Surroundings.
The study of energy and its transformation is known as **THERMODYNAMICS**.

**THERMOCHEMISTRY** is the portion of thermodynamics that pertains to chemical reactions.

The energy possessed by a system is called its **internal energy**.

**Internal energy** of a system is the sum of all kinetic and potential energies of all components of the system.

Note: Most books use $U$ for internal energy. This book uses the older nomenclature $E$.

The **1st Law of Thermodynamics** states that the energy of the universe is constant:

\[ \Delta U_{\text{universe}} = \Delta U_{\text{system}} - \Delta U_{\text{surrounding}} = 0 \]

Energy can be transferred from **system** to **surroundings** (or vice versa) but it can't be created or destroyed:

\[ \Delta U_{\text{system}} = \Delta U_{\text{surrounding}} \]

A more useful form of the first law:

\[ \Delta U_{\text{system}} = q + w \]

Changes involving energy can either be **PHYSICAL** or **CHEMICAL**.

**PHYSICAL Changes**
1. Change in state (No temperature change)
2. Change within a state (Change in temp)

**CHEMICAL Changes** (Reactions)
1. Formation
2. Combustion
3. Neutralization