Additional Short Notes for Chemistry

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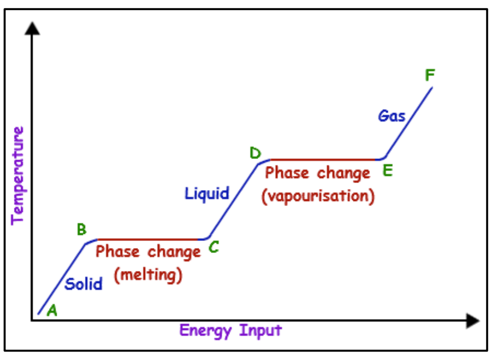
Note: These are things that may or may not come out on the tests, but it's good to have a few aces up your sleeves. Good luck! Look at these lessons one at a time; reading this whole thing in one go is discouraged.

Contents

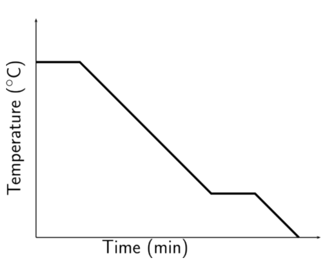
Phase Transitions in Matter

Phase Transition	What it's called	Examples
solid to liquid	melting, fusion	melting ice
solid to gas	sublimation	dry ice, freeze-drying coffee
liquid to solid	freezing	freezing water
liquid to gas	vaporization	steaming
gas to liquid	condensation, liquefaction	formation of dew
gas to solid	deposition	formation of frost and snow

Phase Change Graphs



- When heat is applied to matter, its curve (or graph), when plotted, contain two main observations. It has regions where:
  - *temperature increase (seen in blue, diagonal lines)*
    - this is where the state of matter **stays the same**
  - *there are temperature plateaus where temperature stays constant (red, flat lines)*
    - this is where the **phase transition or phase change occurs** (e.g. melting, vaporization)
- This is also applicable when a substance cools down:



Electronic Configurations

- One can determine an element's group and period using its electronic configuration.
  - Period: Energy Level
    - The highest energy level in an element's configuration refers to its **period** (seen as the rows [horizontal] in the periodic table)

> **Example: Bromine**
> $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
> Period: **4**

  - Group: Valence Electrons
    - The valence electrons (no. of electrons in the highest energy level) of an element refers to its **group** (seen as the columns [vertical] in the periodic table)

> **Example: Bromine**
> $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
> The fourth energy level has a total of 7 electrons.
> Group: **17 or VIIA**
> DISCLAIMER: This method is applicable for representative elements (Groups 1-2, and 13-18). Don't worry though; you'll mostly encounter only these in the tests.

    - This is so because of the Octet Rule. In the above example, only Group A elements (representative elements) follow the Octet Rule, this doesn't work for the Group B elements (transition elements). **Review: What are representative and transition elements? Are there differences between them?**
      - **The Octet Rule:** Elements are stable when they have eight valence electrons.

Chemical Reactions

More Types of Chemical Reactions

- Exothermic and Endothermic Reactions
- Neutralization Reactions

Basic Information

Exothermic Reactions

- A chemical reaction that **releases** heat to proceed**;** this means that heat is considered a **product**
  - Heat goes **out** of the system (*ex-* in the name)
- An example of this involves a **combustion** of fuels, where heat is released, like in fire.
  - Another analogy: when one has an "ex", and he/she is **out** of his/her life

Endothermic Reactions

- A chemical change that absorbs heat energy; heat is considered a **reactant** in this situation.
- An example of this involves melting ice. The ice **absorbs** heat energy for it to melt.

Summary

Endothermic	Exothermic
heat is **absorbed**	heat is **released**
heat is a **reactant**	heat is a **product**
energy must be applied and **work must be done** for it to start	may occur **spontaneously**

Neutralization Reactions

- A chemical reaction that occurs when an **acid and a base combine to form water and a salt**.
- A good example of this involves the common examples for acids and bases: hydrochloric acid and lye (sodium hydroxide:

  $HCl + NaOH \rightarrow NaCl + H_2O$
- This reaction formed water and sodium chloride (also known as table salt).
  - Note that mixing strong acids and strong bases usually form substances with a pH of 7, like water (e.g. the example reaction above).

The Molecular and Empirical Formulas

Empirical Formula

- **The simplest formula for a compound**; this is similar to a fraction being transformed into lowest terms.
- This involves the lowest whole-number ratios between the numbers of elements in a compound.

Molecular Formula

- **The actual formula for a molecule;** this can be an empirical formula or a multiple of it. **Review: What type of chemical bond is present in molecules?**
- **Examples of molecular formulas for C_3HO_2**
  - C_3HO_2, $C_6H_2O_4$, etc.

NOTE: For getting these formulas, values (percentage compositions of the element and/or atomic weights, etc.) are required. However, these will be given in the tests.

Getting the Empirical Formula

Note: Percentage compositions and atomic weights are usually given for this.

- **100g:** Assume 100g of the substance
- **Mass:** Get the masses of each element given 100g of the substance.
- **Moles:** Divide each element's mass by its respective atomic weight. This expresses each element's mass in moles.
- **Smallest:** Divide each mole measure by the smallest acquired mole quantity.
- **Ratio:** Get a ratio based on each element's resulting number.

Example:

Determine the empirical formula of a molecule composed of 30.4% nitrogen and 69.6% oxygen.
(molar masses: N = 14, O = 16)

100g	There is 100g of the substance.	
Mass	Nitrogen: $100 \times 30.4\% = 30.4$ g	Oxygen: $100 \times 69.6\% = 69.6$ g
Moles	$\frac{30.4}{14} \approx \frac{30}{14} \approx 2$ mol N	$\frac{69.6}{16} \approx \frac{70}{16} \approx 4$ mol O
Smallest	$\frac{2 \text{ mol}}{2 \text{ mol}} = 1$	$\frac{4 \text{ mol}}{2 \text{ mol}} = 2$
Ratio	N:O = 1:2 **Therefore, the empirical formula is NO_2.**	

Estimation is important especially when solving for this formula during a test, where no calculators are allowed. (Seen in the "Moles" step)

Getting the Molecular Formula

Note: The molecular weight of the substance is usually given for this, in addition to the requirements for the empirical formula.

- **Empirical:** Get the empirical formula of the substance.
- **Total Mass:** Get the total mass of the substance by adding the respective products of the whole-number ratios to each element's atomic weight. This gives one empirical unit.
- **Divide:** Divide the empirical unit by the given molecular weight.
- **Multiply:** Multiply each element's coefficient by the given quotient to get the molecular formula.

Example:

Determine the molecular formula of a molecule composed of 30.4% nitrogen and 69.6% oxygen and has a molecular weight of about 138 g/mol. (atomic weights: N = 14, O = 16).

Empirical	N:O = 1:2 **The empirical formula is NO_2.** (As seen in the previous example).
Total Mass	$1(14)+2(16)=14+32$ $=46\text{ g/mol}$ One mole of the substance weight 46 grams. (This is one empirical unit)
Divide	$\frac{138}{46}=3$ There are there empirical units in one molecule of the substance.
Multiply	$NO_2 \rightarrow N_3O_6$ **The molecular formula is N_3O_6** since there are 3 empirical units in a molecule.

Basic Stoichiometry

The Basic Concept

Stoichiometry is branch of science that deals with the quantitative relationships between reactants and products in a chemical reaction.

In English:

How much of the reactants will transform into these products? (and vice-versa)
One should remember this question when performing calculations for stoichiometry.

Example:
In this reaction: $2H_2 + O_2 \rightarrow 2H_2O$, it can be determined that:

- **2 moles** of hydrogen and **1 mole** of oxygen will transform into **2 moles** of water, and/or;
- **2 moles** of water will be produced if **2 moles** of hydrogen and **1 mole** of oxygen are used as reactants
  - Note that the number of moles in the relationship can be seen in the **coefficients** of each component in the reaction.

Performing Stoichiometric Calculations

- Questions involving stoichiometry usually ask how much of one reactant produces a given amount of product, and vice-versa.
- There are two main measures that are used in these questions: **mole, and mass.**
  - **Review: One mole of an element is equal to its molar mass in mass, and vice versa.**

Solving Problems in Basic Stoichiometry

How to Practice Solving These Problems

- **Balance:** Check and make sure that the equation is balanced. **Review: How does one balance chemical equations?**
- **Table:** It's a good start to use a table. :)
- **Organize:** Organize in it how many moles and mass (given one mole) are involved in the reaction.
- **Ratio:** Write the ratio of the involved elements' mole measures.
- **Solve:** Given two ratios, (one will be unknown: what we're solving for), solve for the unknown quantity. **Get ready to convert at any point during solving.** **Review: How to convert?**

Example 1:
Given this reaction: $2H_2 + O_2 \rightarrow 2H_2O$, how many moles of water is produced by 5 moles of oxygen?
(molar masses: H = 14, O = 16)

<table>
<tr><td>Balance</td><td>This equation is balanced.</td></tr>
<tr><td>Table</td><td>
<table>
<tr><td></td><td>$2H_2$</td><td>O_2</td><td>$2H_2O$</td></tr>
<tr><td>Moles</td><td></td><td></td><td></td></tr>
<tr><td>Molar mass</td><td></td><td></td><td></td></tr>
</table>
</td></tr>
<tr><td>Organize</td><td>
<table>
<tr><td></td><td>$2H_2$</td><td>O_2</td><td>$2H_2O$</td></tr>
<tr><td>Moles</td><td>2</td><td>1</td><td>2</td></tr>
<tr><td>Molar mass</td><td>2(1)
= 2 g/mol</td><td>2(16)
= 32 g/mol</td><td>2(1) + 16
= 18 g/mol</td></tr>
</table>
</td></tr>
<tr><td>Ratio</td><td>$O_2 : H_2O = 1:2$
$1:2 = 5:x$
Given this ratio, x represents the number of moles of water produced.</td></tr>
<tr><td>Solve</td><td>$1:2 = 5:x$
$x = 10$
Solve for x in this ratio, this gives x = 10.
Therefore, 10 moles of water are produced.</td></tr>
</table>

Example 2:
Given this reaction: $NaClO_3 \rightarrow NaCl + O_2$

(1) 12.00 moles of $NaClO_3$ will produce how many grams of O_2?
(2) How many grams of NaCl are produced when 80.0 grams of O_2 are produced?

(molar masses: Na = 23, Cl = 35.45, O = 16)

<table>
<tr><td>Balance</td><td colspan="2">This equation is not balanced.
The balanced equation is:
$2NaClO_3 \rightarrow 2NaCl + 3O_2$</td></tr>
<tr><td>Table</td><td colspan="2">
<table>
<tr><td></td><td>$2NaClO_3$</td><td>$2NaCl$</td><td>$3O_2$</td></tr>
<tr><td>Moles</td><td></td><td></td><td></td></tr>
<tr><td>Molar mass</td><td></td><td></td><td></td></tr>
</table>
</td></tr>
<tr><td>Organize</td><td colspan="2">
<table>
<tr><td></td><td>$2NaClO_3$</td><td>$2NaCl$</td><td>$3O_2$</td></tr>
<tr><td>Moles</td><td>2</td><td>2</td><td>3</td></tr>
<tr><td>Molar mass</td><td>106.44
g/mol</td><td>58.44
g/mol</td><td>32 g/mol</td></tr>
</table>
</td></tr>
<tr><td rowspan="2">Ratio</td><td>For (1)</td><td>For (2)</td></tr>
<tr><td>$NaClO_3 : O_2 = 2:3$
$2:3 = 12:x$</td><td>$NaCl : O_2 = 2:3$
$2:3 = x:80\text{ g}$
80 g needs to be converted to moles of O_2 for the mole ratio.</td></tr>
<tr><td>Solve</td><td>$2:3 = 12:x$
$x = \frac{(12)(3)}{2} = 18\text{ mol}$</td><td>Conversion:
$80\text{ g}\left(\frac{1\text{ mol}}{32\text{ g}}\right) = 2.5\text{ mol}$
With this, the new mole ratio is:
$2:3 = x:2.5$</td></tr>
<tr><td rowspan="2">Solve (cont.)</td><td>There are 18 moles of O_2 produced. This is expressed in grams as:
$18\text{ mol}\left(\frac{32\text{ g}}{1\text{ mol}}\right) = 576\text{ g}$</td><td>Using the ratio
$2:3 = x:2.5$, solve for x.
$x = \frac{(2.5)(2)}{3} = \frac{5}{3}$
$= 1.6667\text{ mol}$
1.6667 mol of NaCl is produced. This is expressed in grams as:
$1.6667\text{ mol}\left(\frac{58.44\text{ g}}{1\text{ mol}}\right) = 97.4\text{ g}$</td></tr>
<tr><td>Final Answer:
576 g of O_2 are produced.</td><td>Final Answer:
97.4 g of NaCl are produced.</td></tr>
</table>

Your skills in conversion and ratios will be crucial in solving these types of problems, so practicing with these is essential.

Molarity and Molality

Basic Information

- Molarity and molality are both **measures** of solutions' **concentrations**. **Review: What are concentrations?**
- These two measurements both express a solution's concentration, but use different aspects of the solution. These involve the **solute** and the **solvent** in the solution.
  - **Solute - the substance dissolved**
  - **Solvent - the substance that dissolves**

Molarity

- Also known as the molar concentration of a solution
- A measurement of **how much solute there is in a solution;** this is expressed in mol/L or a unit called *molar* (M).
- To put it simply: it is the **amount of substance in a given volume**

What is measured	Standard Units	Formula for molarity
solute	moles (mol)	$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$
solution	liter (L)	

Molality

- Also known as the molal concentration of a solution
- A measurement of **how much solute there is in terms of its solvent;** this is expressed in mol/kg or a unit called *molal* (m).

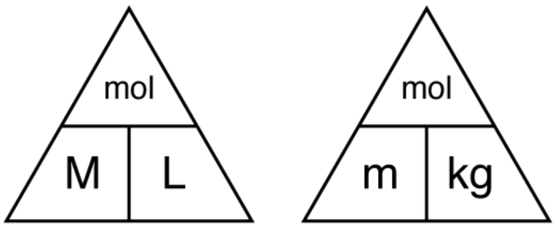
What is measured	Standard Units	Formula for molarity
solute	moles (mol)	$\text{molality} = \frac{\text{moles of solute}}{\text{kg of solvent}}$
solvent	kilograms (kg)	

Summary

	Molarity	Molality
What they are	measurements of a solution's concentration	
What they need to measure	moles of solute; liters of solution	moles of solute; kg of solvent
Units of measurement	molar (M) or $\frac{\text{mol}}{\text{L}}$	molal (m) or $\frac{\text{mol}}{\text{kg}}$
What one unit is	*one mole of a substance in one liter of solution*	*one mole of a solute dissolved by one kg of solvent.*

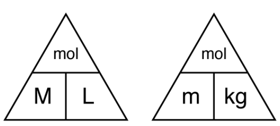
Solving for Basic Molarity and Molality

- **Identify:** Which is the *unknown* in the situation?
- **Convert:** Check and make sure that measurements are in their standard units (See Basic Information)
  - **Review: One mole of an element is equal to its molar mass in mass, and vice versa.**
- **Molar Mass:** Get the molar mass of the compound.
- **Triangles:** Using these triangle figures much like the speed-distance-time ones can greatly help in solving for these measurements.

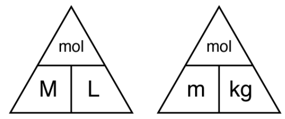


  - In these figures:
    - M = **molarity** of a solution in mol/L
    - m = **molality** of a solution in mol/kg
    - L = volume of the **solution in liters** (L)
    - kg = mass of the solvent in kilograms (kg)
- **Solve:** Using the triangles, solve for missing quantities and convert if necessary.

Example 1:
How many grams of potassium carbonate (K_2CO_3) are needed to make 200 mL of a 2.5 M solution?
(molar masses: K = 39.1, C = 12, O = 16)

Identify	The identified unkown is the moles of solute needed (mol).
Convert	200 mL must be converted to liters: $200\text{ mL}\left(\frac{1\text{ L}}{1000\text{ mL}}\right) = 0.2\text{ L}$
Molar Mass	$2(39.1)+12+3(16)=138.2\text{ g/mol}$ There are 138 g in one mole of K_2CO_3.
Triangles	 Based on the triangles, the one with the M (for molarity) must be used. Based on this triangle, getting the moles of solute needed involves **multiplying** the molarity and the volume. $\text{mol} = \text{M} \times \text{L}$
Solve	$\text{moles of solute} = 2.5\text{ M} \times 0.2\text{ L} = 0.5\text{ mol}$ Converting this to grams gives: $0.5\text{ mol}\left(\frac{138.2\text{ g}}{1\text{ mol}}\right) = 69.1\text{ g}$ **Therefore, 69.1 g of potassium carbonate is needed.**

Example 2:
Determine the molality of a solution with 100 g of sucrose, with a molar mass of 342.3 g/mol, dissolved in 1.50 kg of water.

Identify	The identified unkown is the molality of the solution.
Convert	100 g must be converted to moles: $100\text{ g}\left(\frac{1\text{ mol}}{342.3\text{ mL}}\right) = 0.292\text{ mol}$
Molar Mass	The molar mass has already been given in the problem.
Triangles	 Based on the triangles, the one with the m (for molality) must be used. Based on this triangle, getting the molality of the solution involves **dividing** the moles of solvent and the mass of the solute. $\text{m} = \frac{\text{mol}}{\text{kg}}$
Solve	$\text{molality} = \frac{0.292\text{ mol}}{1.50\text{ kg}} = 0.195\text{ m}$ **Therefore, the molality of the solution is 0.195 m.**

Other Stuff

Acids and Bases

- **This must be remembered:**
  - Acids are associated with **hydrogren** ions. [H^+ ions].
  - Bases are associated with **hydroxide** ions. [$(OH)^-$ ions].
- **Definitions:** Acids and bases have different definitions in chemistry. These definitions may be asked for in the tests as knowledge questions.

Definitions of Acids and Bases	**Acid**	**Base**
Arrhneius	A substance that yields H^+ ions in an aqueous solution	A substance that yields OH^- ions in an aqueous solution
Brønsted–Lowry (proton-centered)	A proton donor	A proton acceptor
Lewis (electron-centered)	An electron pair acceptor	An electron pair donor

STP (Standard Temperature and Pressure)

- This set of conditions is the most commonly used when performing calculations involving gases.

Temperature	**Pressure**	**Volume of one Mole of Gas**
0°C = 273.15 K	1 atm = 760 torr	22.4 L

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