Chemistry 11 - Week 4: May 4 – May 8

Anticipated time required: 3 hours

New learning objective: **An Introduction to Chemical Reactions**

Goals to be completed:

Check out youtube for a variety of really cool chemical reactions!

Here's a good place to get started:

https://www.youtube.com/watch?v=KNPoBoUt-HM

1. Laws of chemical reactions
2. Types of chemical reactions
3. Endothermic and exothermic reactions
4. Balancing chemical reactions
5. Practice on balancing
6. Formal unit assignment
7. Mole Unit Summary Assignment

This PDF package contains several notes, examples and practice problems. The only formal portion that you are required to submit is the section titled "Formal Assignment to be Submitted". This can be sent to Charlie.feht@yesnet.yk.ca either as a scanned and uploaded PDF attachment to email, or as a jpeg image file. Midterm assignments will be scored and sent back to you as I receive them.

Upcoming next week:

Stoichiometry and crossing the mole bridge

Section 1: The laws of chemical reactions

A chemical reaction shows the chemicals used up in a chemical reaction. This is described as:

Reactants → Products

Where reactants are the existing chemicals, products are the existing new chemicals, and "→" indicates the passage of time.

Chemical reactions can be written in words or symbols.

Ex. Sodium and chlorine react to form sodium chloride
Ex. $2Na + Cl_2 \rightarrow 2NaCl$

We know that a chemical reaction has taken place if we observe one or more of the following:

1. Temperature change has occurred
2. Colour of the system changes
3. A precipitate forms
4. A gas is produced (not a reliable measure on its own)

Chemical reactions can take place in an **open or a closed system**. In an **open system**, the reaction can be influenced by surrounding chemicals, as things are able to enter and leave the system (i.e. burning an open flame). In a closed system, nothing can enter or leave the system (i.e. boiling water with a lid covering a pot).

During all chemical reactions, four conservation laws must be followed:

1. The law of conservation of mass
    a. When I react 50 g of of a reactant, I should produce 50 g of product
        i. Note – this may not be measured in a lab experiment, as many labs are performed in and open system, and some mass may escape from our measurements. That does not mean the mass is not produced.

2. Law of constant proportions
    a. Otherwise known as balancing equations. There must be the same amount of atoms on the reactant and product side
        i. $2Na + Cl_2 \rightarrow 2NaCl$ contains 2 sodium and 2 chlorine atoms on each side

3. Electrical charge
    a. Chemical reactions do not gain or lose electrical charge in a chemical reaction

4. Law of conservation of energy
    a. Energy cannot be created or destroyed, it only changes form.
        i. A ball on a perch has high potential energy, but as it is dropped it loses potential energy, yet gains kinetic energy making the system completely equal to the point where no energy is lost in the system.

Helpful summary video for reaction types: https://www.youtube.com/watch?v=z0187f8gqfU

Section 2: Types of chemical reactions

1. Synthesis reaction: A + B → AB
    a. Occurs when 2 elements/molecules react to form one compound
        i. $H_2 + 2Cl \rightarrow 2HCl$
2. Decomposition reaction: AB → A + B
    a. Occurs when one compound reacts to form 2 elements
        i. $2Ag_2O \rightarrow 4Ag + 0_2$
3. Single replacement reaction: AB + C → CB + A
    a. When one element replaces the position of a different element comprising a molecule. A lone metal will replace a metal in a molecule, a lone nonmetal will replace a nonmetal in a molecule.
        i. $CuCl_2 + Fe \rightarrow FeCl_2 + Cu$
4. Double replacement reaction: AB + CD → CB + AD
    a. The metals of two compounds will swap places and form new molecules.
        i. $Ba(NO_3)_2 + Na_2SO_4 \rightarrow 2NaNO_3 + BaSO_4$
    b. A neutralization reaction can also occur here, when the reactants are an acid and a base. An acid and a base will always react to form water, and one other ionic compound (otherwise known as a salt).
        i. Acid + base → salt + water
        ii. $HCl + NaOH \rightarrow NaCl + H(OH)$
5. Combustions: Hydrocarbon + $O_2 \rightarrow CO_2 + H_2O$
    a. Any hydrocarbon (molecule made of carbon and hydrogen) will burn in the presence of oxygen to produce carbon dioxide and water
        i. $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

Please be aware, not all compounds conform to these 5 reaction types. Some violate the expected pattern of behaviour, such as in the case of sulfur dioxide and water reacting to form sulfurous acid. $SO_2 + H_2O \rightarrow H_2SO_3$

Practice Problems.

Label the reaction type below:

1. $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$

2. $NaCl(aq)\ AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

3. $H_2O(l) + SO_3(g) \rightarrow H_2SO_4(aq)$

4. $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

5. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

6. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

Section 3: Endothermic and exothermic reactions

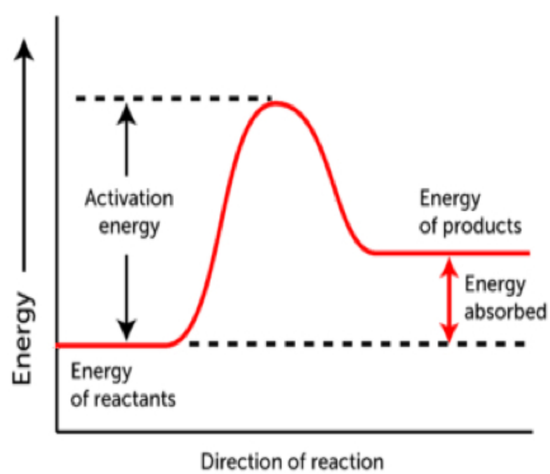
In order to undergo chemical reactions, chemical bonds in the reactants must be broken and new ones formed in the products. This requires energy; either consumed in the breaking of bonds, or released in the formation of bonds.

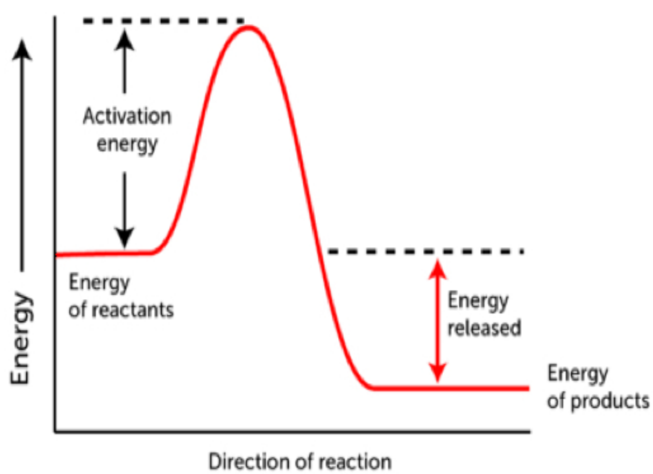
The net change in energy required during the chemical reaction determines whether or not the reaction is endothermic or exothermic, and is measured in kilojoules (kJ).

1. Exothermic: When heat is given off (exits the system)
    a. In this scenario, less energy is used to break the bonds than there is produced when the new ones form. Energy is therefore given off as heat and the surroundings of the system feel warmer.
    b. In a chemical reaction this is represented by a kJ value on the product side
        i. $CH_4 + 2O_2 \rightarrow CO_2 + H_2)$ **+ 891 kJ**

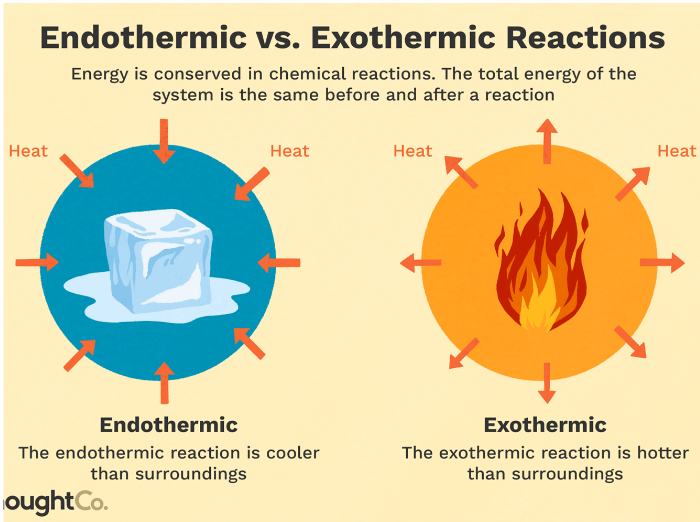
2. Endothermic: When heat is added (heat enters the system)
    a. In this scenario, more energy is required to break the bonds than there is produced when new ones form. Energy is therefore absorbed in the system to complete the reaction and the surroundings will feel cooler.
    b. In a chemical reaction this is represented by a kJ value on the reactant side
        i. $KClO_3$ + **41.4 kJ** $\rightarrow K^+ + ClO_3^-$

Practice: Can you label the reaction type below as endothermic or exothermic based on the diagram of energy values?





Helpful summary video on thermochemistry (enthalpy, exothermic, endothermic): https://www.youtube.com/watch?v=ZVhJ4TO8a-M



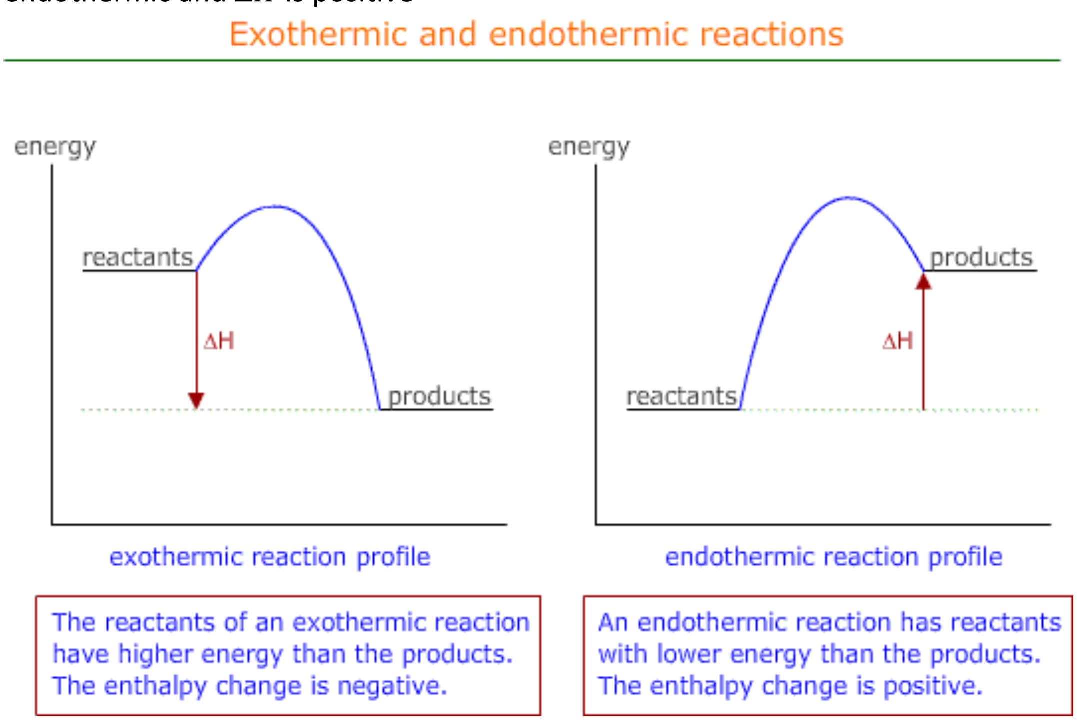
Enthalpy is a measure of <u>the change in heat in a system</u>. This is measured by the symbol ΔH.

$$\Delta H = \text{heat of the product} - \text{heat of the reactant}$$

For this reason,

- If the product has lower energy than the reactant (energy given off) the reaction is exothermic and ΔH is negative.

- If the product has higher energy than the reactant (energy absorbed) the reaction is endothermic and ΔH is positive



Play around with the following simulation to help solidify your balancing skills!

https://phet.colorado.edu/sims/html/balancing-chemical-equations/latest/balancing-chemical-equations_en.html

Section 4: Balancing Chemical Reactions

Chemical reactions are like recipes in that the quantity and types of ingredients, or **reactants**, can be related to the quantity and type of cooked food, or **product(s)** formed. Balancing chemical reactions then allows one to determine stoichiometry calculations by understanding the ratio between reactants and/or products. This worksheet includes some rules and guidelines to help you balance chemical equations.

Rules

1.) The formulas of the reactants and products **cannot** be changed, do not alter subscripts or charges.
2.) The **only** numbers that can be changed are the numbers indicating how many molecules or atoms, which are called **coefficients**.
3.) A coefficient is assumed to be **one** if there is not a number in front of the molecule or atom.
4.) In order to be balanced, there must be an equal number of each type of atom on both the reactant and product side of the reaction.
5.) It is generally required that the coefficients are **whole numbers**.

Guidelines

1.) In reactions dealing solely with ions, one can leave the polyatomic ions as groups for ease of balancing.
2.) In reactions dealing with only ions and water, water can be considered as a combination of a **hydrogen ion and hydroxide ion**.
3.) If given a reaction with polyatomic ions that are broken down, one **cannot** leave the polyatomic ions as groups.
4.) "Atom accounting" makes this easier by using a table, and is detailed in the following guideline points and examples.
5.) Start with all coefficients of one and total the number of each type of atom or species.
6.) The more atoms in a given molecule, the **larger** the effect it has on balancing, so begin with these.
7.) End with molecules or atoms that consist of only one type, since the number can be changed **independently** of the other atom types.
8.) If a coefficient comes out to a fraction, multiply all coefficients by the fraction denominator to result in all whole number coefficients.

Practice

Chapter 7 Worksheet #1

Balancing Chemical Equations

Balance the equations below:

1) ____ N_2 + ____ H_2 → ____ NH_3

2) ____ $KClO_3$ → ____ KCl + ____ O_2

3) ____ $NaCl$ + ____ F_2 → ____ NaF + ____ Cl_2

4) ____ H_2 + ____ O_2 → ____ H_2O

5) ____ $Pb(OH)_2$ + ____ HCl → ____ H_2O + ____ $PbCl_2$

6) ____ $AlBr_3$ + ____ K_2SO_4 → ____ KBr + ____ $Al_2(SO_4)_3$

7) ____ CH_4 + ____ O_2 → ____ CO_2 + ____ H_2O

8) ____ C_3H_8 + ____ O_2 → ____ CO_2 + ____ H_2O

9) ____ C_8H_{18} + ____ O_2 → ____ CO_2 + ____ H_2O

10) ____ $FeCl_3$ + ____ $NaOH$ → ____ $Fe(OH)_3$ + ____$NaCl$

11) ____ P + ____O_2 → ____P_2O_5

12) ____ Na + ____ H_2O → ____ $NaOH$ + ____H_2

13) ____ Ag_2O → ____ Ag + ____O_2

14) ____ S_8 + ____O_2 → ____ SO_3

15) ____ CO_2 + ____ H_2O → ____ $C_6H_{12}O_6$ + ____O_2

16) ____ K + ____ $MgBr_2$ → ____ KBr + ____ Mg

17) ____ HCl + ____ $CaCO_3$ → ____ $CaCl_2$ + ____H_2O + ____ CO_2

18) ____ HNO_3 + ____ $NaHCO_3$ → ____ $NaNO_3$ + ____ H_2O + ____ CO_2

19) ____ H_2O + ____ O_2 → ____ H_2O_2

20) ____ $NaBr$ + ____ CaF_2 → ____ NaF + ____ $CaBr_2$

21) ____ H_2SO_4 + ____ $NaNO_2$ → ____ HNO_2 + ____ Na_2SO_4

Practice

Word Equations

Write the word equations below as chemical equations and balance:

1) Zinc and lead (II) nitrate react to form zinc nitrate and lead.

2) Aluminum bromide and chlorine gas react to form aluminum chloride and bromine gas.

3) Sodium phosphate and calcium chloride react to form calcium phosphate and sodium chloride.

4) Potassium metal and chlorine gas combine to form potassium chloride.

5) Aluminum and hydrochloric acid react to form aluminum chloride and hydrogen gas.

6) Calcium hydroxide and phosphoric acid react to form calcium phosphate and water.

7) Copper and sulfuric acid react to form copper (II) sulfate and water and sulfur dioxide.

8) Hydrogen gas and nitrogen monoxide react to form water and nitrogen gas.

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To Submit

Chemistry 11 ~~Extra~~ Assignment ~~#1~~

1. Balance the following chemical equations:
   a. $\_\ Be_2C + \_\ H_2O \rightarrow \_\ Be(OH)_2 + \_\ CH_4$

   b. $\_\ NpF_3 + \_\ O_2 + \_\ HF \rightarrow \_\ NpF_4 + \_\ H_2O$

   c. $\_\ NO_2 + \_\ H_2O \rightarrow \_\ HNO_3 + \_\ NO$

   d. $\_\ Cu + \_\ HNO_3 \rightarrow \_\ Cu(NO_3)_2 + \_\ NO + \_\ H_2O$

2. Balance the following word equations:
   a. Silver nitrate and copper are reacted to form solid silver and aqueous copper (II) nitrate.

   b. Sodium chloride and sulfuric acid react to produce hydrochloric acid and sodium sulfate

   c. Phosphoric acid is neutralized in potassium hydroxide to yield water and potassium phosphate

   d. Copper (II) Iodide reacts with bromine to produce copper (II) bromide and iodine.

3. Identify the following reaction types:
   a. Aluminum reacts with sulfur to produce aluminum sulfide

   b. Nitric acid reacts with strontium hydroxide to produce strontium nitrate and water

   c. Dicarbon dihydride burns in oxygen to produce carbon dioxide and water

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d. Iron (II) chloride and potassium sulfide react to form ion (II) sulfide and potassium chloride

e. Hydrogen iodide breaks down to yield hydrogen gas and iodine.

4. Complete, identify and balance the following reactions:
   a. Sodium phosphate + Calcium hydroxide →

   b. Chlorine gas + Calcium bromide →

   c. HCl + KOH →

   d. $CH_4 + O_2 \rightarrow$

5. Determine if the following reactions are endothermic or exothermic
   a. $C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O + 5638\ kJ$

   b. $KClO_3 + 41.4\ kJ \rightarrow K^+ + ClO^-$

   c. $\Delta H = -432\ kJ$

   d. $\Delta H = +891\ kJ$

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To Submit

Mole Concept Summary Assignment

Multiple Choice:

1. The number of molecules in a mole of any molecular substance compared to the number of atoms in one mole of any element is

1) always larger;
2) always less;
3) sometimes larger and sometimes smaller;
4) always the same;

2. The molar mass of an element is always equal to:

1) The atomic weight
2) 6.02×10^{23}
3) The atomic mass
4) Avogadro's number

3. What is the molar mass of ammonium sulfate?

1) 132.l g/mol.
2) 114 g/mol
3) 118 g/mol
4) 131.1 g/mol

4.There are 6.02×10^{23} water molecules in a mole of water. What is the mass of 3.01×10^{23} molecules of water?

1) 0.50 grams
2) 18.0 grams
3) 9.00 grams
4) 27.0 grams
5) 36.0 grams

5. A gas has a density of 0.717 g/l at STP. What is the mass of one mole of this gas?

1) 8.0 g.
2) 16.1 g.
3) 28.1 g.
4) 4.0 g.

6. How much space will 5.5×10^{20} molecules of oxygen occupy at STP conditions?

1) 2.1×10^{-2} 1iters
2) 6.8×10^{-1} liters
3) 1.5×10^{3} liters
4) 2.4×10^{4} liters

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7. What is not a proper way to represent molarity?

1) M; mol/L; Concentration
2) Molar; mol/L, C
3) [square brackets]; M; Molar
4) [square brackets] L/mol; Concentration

8. What is the molar volume of 1 mol of O_2 gas at STP?

1) 22.4 g
2) 22.4 L
3) 32 g
4) 32 L

9. What mass of solid calcium chloride, $CaCl_2$, is dissolved in one liter of water to prepare a 0.40 molar solution?

1) 4.40 grams
2) 14.8 grams
3) 44.4 grams
4) 55.5 grams
5) 111.0 grams

10. If 328 grams of nickel hydrogen carbonate are dissolved to make 2.7 liters of solution, what is the molar concentration of this solution?

1) 0.67 M
2) 1.33 M
3) 2.80 M
4) 6.70 M

Mole Concept Conversion Problems

1.) What volume at STP is occupied by 8.65 molecules of SO_2 ?

2.) How many grams are there of carbon dioxide are there in a 30.57 L sample at S.T.P.

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3.) What mass of NaOH is contained in 3.50 L of 0.200 M NaOH?

4.) 28.56 g of KCl is dissolved into water to make 300.00 mL of solution. What is the molar concentration of the solution?

Percent Composition and Empirical/Molecular formula problems

1. Calculate the Percent composition by mass of $K_3Fe(CN)_6$

2. A compound was found to contain 43.7 g of phosphorous and 56.3 grams of oxygen. What is the empirical formula for the compound?

3. What is the empirical formula of a compound containing 77.9% I and 22.1% O?

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4. The empirical formula of a compound is SiH_3. If 0.0275 mol of a compound has a mass of 1.71 g, what is the compounds molecular formula?

Mixing and Dilution Problems

1.) 45.00 ml of a 1.57 M solution of NaOH is diluted by adding 125.00 ml of water. What is the concentration of the diluted solution?

2.) How many mL of a 12.5 M solution of HCl does it take to make 4.5 L of a 0.500 M solution of HCl?

3.) 235.0 ml of a 1.67 M solution of NaCl is mixed with 550.0 ml of 1.33M solution of NaCl what is the Molarity of the mixture?