

I. GO OVER SOLUTIONS FOR GROUP PROBLEM 6

II. REVIEW FOR EXAM 2 (TUESDAY, APRIL 8.....2:30-3:45)

I. Solutions for Group Problem 6

1) Arrange the following kinds of electromagnetic radiation in order of increasing wavelength:

- Infrared
- Microwaves
- Green
- X-rays
- Orange light
- UV radiation

Answer: X-rays, UV radiation, Green, Orange, Infrared, Microwaves. Out of these wavelengths, X-rays have the smallest wavelength and Microwaves have the biggest wavelength. If you reverse the answer's order, you will get the order of increasing frequency. This is because wavelength and frequency are inversely related. Also, energy and frequency are directly related, so the frequency order is the same as the energy order.

(See Fig. 6.4 on page 219 in Brown/LeMay book for picture of electromagnetic spectrum)

*To remember visible light colors and their order, use the pneumonic device: **ROYGBV** or red, orange, yellow, green, blue, and violet. These colors are in the order of increasing frequency and energy, but decreasing wavelength.

2) Answer the following questions about radiation that has wavelength 236 nm:

- a. What is the frequency of the radiation?
- b. What kind of radiation is it?
- c. What is the energy of a photon of this wavelength?
- d. How many kJ of energy does 1 mol of photons have?

Answers: a. Since we need to use the formula $c = \lambda\nu$, where $c = 2.998 \times 10^8$ m/s (speed of light), λ = wavelength, and ν = frequency, we first need to convert 236 nm into meters:

$$236\text{nm} \times \frac{10^{-9} \text{ m}}{1\text{nm}} = 2.36 \times 10^{-7} \text{m}$$

Now we use $c = \lambda\nu$, but we need to find ν , so:

$$\nu = \frac{c}{\lambda}$$

$$\nu = \frac{2.998 \times 10^8 \text{m/s}}{2.36 \times 10^{-7} \text{m}} = 1.27 \times 10^{-7} \text{s}^{-1}$$

*Note: The unit for frequency is $\text{s}^{-1} = \frac{1}{\text{s}} = 1\text{H}$

b. UV radiation. Look at the electromagnetic spectrum and the powers of ten. $2.36 \times 10^{-7} \text{m}$ (wavelength) and $1.27 \times 10^7 \text{s}^{-1}$ (frequency) both fall under the UV radiation part of the spectrum.

c. To find the energy of a photon, we use the formula $E = h \nu$, where E is the energy (in Joules) of a single photon, h is a constant ($6.626 \times 10^{-34} \text{ J} \cdot \text{s}$), and ν is the frequency (in this case, we calculated it to be $1.27 \times 10^7 \text{ s}^{-1}$):

$$E = h \nu$$

$$E = (1.27 \times 10^7 \text{ s}^{-1})(6.626 \times 10^{-34} \text{ J} \cdot \text{s})$$

$$E = 8.42 \times 10^{-19} \text{ J}$$

d. To find the energy (in kJ) produced by 1 mol of photons, we need Avogadro's Number because there are 6.022×10^{23} photons in one mol of photons:

$$\frac{8.42 \times 10^{-19} \text{ J}}{\text{photon}} \times \frac{6.022 \times 10^{23} \text{ photons}}{1 \text{ mol of photons}} = \frac{5.07 \times 10^5 \text{ J}}{\text{mol}}$$

remember to convert J to kJ: $5.07 \times 10^5 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 5.07 \times 10^2 \text{ kJ}$ or 507 kJ

II. Review for Exam 2

-Exam 2 will cover material from homework assignments 6 through 8, this includes 6 and 8. In other words the last part of Ch. 4, all of Ch. 5, and the first part of Ch. 6.

Key Concepts to Study:

-*Concentration of solutions (Molarity)*: figuring out concentrations of certain ions in solutions, titration problems (using stoichiometry).

-*Redox Rxns*: knowing how to figure out oxidation numbers for atoms, which ones are being reduced or oxidized, which ones are the reducing agent or oxidizing agent.

-*Writing net ionic equations*: redox rxns, weak acid + strong base vs. strong acid + strong base.

You need to remember the strong acids: HCl, HBr, H_2SO_4 , HI, HClO_3 , HClO_4 ,



strong bases: LiOH, KOH, RbOH, CsOH, $\text{Ca}(\text{OH})_2$,



-*Heat Rxns*: system vs. surrounding, enthalpy of rxn, standard enthalpy of formation, Hess' Law

Concentration/Molarity

Sample Problem What is the concentration of a solution that was made by adding 1.95g of Na_2SO_4 to 250. mL of water? (f.w. of Na_2SO_4 is $142.05\text{g}\cdot\text{mol}^{-1}$)

1) First, you have to realize that this is a molarity question because it's asking about a concentration of a particular molecule. You also have to know the formula for molarity (M):

$$M = \frac{\text{mol of solute}}{\text{L of solution}}$$

2) We need to know how many moles are in 1.95g of Na_2SO_4 by using the given f.w. of Na_2SO_4 :

$$1.95\text{g} \times \frac{1\text{mol}}{142.05\text{g}} = 0.01373 \text{ mol Na}_2\text{SO}_4$$

3) Now we need to convert 250. mL of water into L to use the molarity equation:

$$250. \text{ mL} \times \frac{1\text{L}}{1000\text{mL}} = 0.250 \text{ L}$$

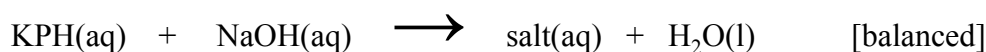
4) Plug the numbers into the molarity equation to solve:

$$M = \frac{0.01373 \text{ mol Na}_2\text{SO}_4}{0.250 \text{ L}} = 0.0549 \text{ mol/L}$$

Sample Problem (Titration) Solid sodium hydroxide absorbs moisture from the air, so it is difficult to weigh accurately to make a solution of known concentration. Potassium hydrogen phthalate, or KHP (m.w. 204.23), is monoprotic, can be weighed accurately, and is often used to standardize solutions of bases. 0.05182 grams of KHP are placed in a flask beneath a buret filled with NaOH solution of unknown concentration. If 23.4 mL of NaOH solution are required to exactly neutralize the KHP solution, what is the concentration of the NaOH?

Solution: We need to find the concentration or **molarity (M)** of NaOH. Since we already have the volume ($23.4\text{mL} = 0.0234\text{L}$), we just need the amount in moles of NaOH. So,

1) Write the acid-base reaction:



2) Next, use the grams and molar mass (or m.w.) of KHP to find the moles of KHP:

$$\frac{0.05182\text{g KHP}}{204.23\text{g/mol}} = 2.537 \times 10^{-4} \text{ mol KHP}$$

3) Use the moles of KHP and coefficients of the acid-base reaction to find moles of NaOH:

$$2.537 \times 10^{-4} \text{ mol KHP} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol KHP}} = 2.537 \times 10^{-4} \text{ mol NaOH}$$

4) Place the grams of NaOH over the volume (**in Liters**) to get **molarity**:

$$\frac{2.537 \times 10^{-4} \text{ mol of NaOH}}{0.0234 \text{ L of NaOH}} = \mathbf{0.0109 \text{ M of NaOH}}$$

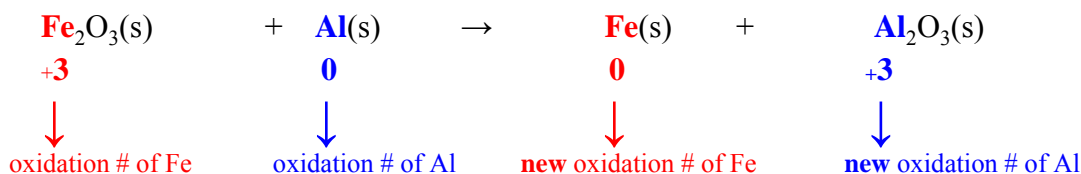
Redox Rxns:

Look at Lecture 17 notes, page 7 for how to get oxidation #s for elements, ions, and elements in a compound

-**electrons** (which have a negative charge) are transferred between two atoms. For every **reduction reaction**, there is an **oxidation reaction** because one atom gives an electron away and another one must pick it up.

-when an acceptor picks up electrons, it is **reduced** (think less in terms of the oxidation number). In contrast, the electron donor is **oxidized**, where its oxidation number increases.

Ex: Iron (III) oxide + aluminum



- as the **iron oxide** and **aluminum** react, aluminum gives iron 3 electrons. Iron is being reduced and aluminum is being oxidized. The species that donates electrons is called the **reducing agent** (in this case it is Al) and the species that receives the electrons is the **oxidizing agent** (in this case, Fe₂O₃).

*when naming the reducing or oxidizing agents, you have to name the whole compound in which the element being reduced or oxidized

Writing Net Ionic Equations (Acid-Base Rxns)

- know the strong acids and bases mentioned above
- strong acids and bases ionize almost completely, which leads to a canceling of some ions in net ionic equation)
- weak acids and bases do not completely ionize, leaving molecules in the net ionic equation. For examples, look at Lecture 17 slides, page 9.

Heat Exchange Rxns

- **kinetic energy** is the energy of motion. Anything that is moving has kinetic energy. The higher the object's velocity (or speed), the more kinetic energy it has.
- potential energy** is stored energy that can be transformed into kinetic energy. For example, a bike standing motionless at the top of a hill has lots of potential energy but no kinetic energy. Once the bike begins to roll downhill, its potential energy decreases as its kinetic energy increases. (See figure 5.2 in book, bottom of page 168)

a **system** is the defined portion of matter (imagine the matter is in a 3-D box where heat can be transferred but not matter) under study in thermodynamics. In contrast, the **surroundings** consist of everything that is not the system, or everything outside the system.

- energy, or heat, can be transferred from a system to its surroundings and the surroundings can also transfer heat to a system. The terms **endothermic** and **exothermic** are used to describe the direction of the heat exchange during a reaction.
- from the perspective of a system:
 - in an **endothermic** reaction-----heat is received by the system by the surroundings
 - in an **exothermic** reaction-----heat is given by the system to the surroundings

Sample Problem

An aluminum block is placed into a beaker of water under certain conditions...

Given Info:

Initial temperature of an aluminum block = 90.00°C
Mass of Al block = 5.00g
Mass of water = 100.0g
Temperature of water before = 23.00 °C
Temperature of both after = 23.71°C
C of water = 4.184 J/gK

Question:

What is the heat capacity of Al?

- 1) We need to find C of Al. Before this, we need to find the ΔT of Al:

$$\Delta T_{Al} = 23.71^{\circ}\text{C} - 90.00^{\circ}\text{C} = -66.29 \text{ K}$$

2) So far, we have:

$$q_{Al} = 5.00\text{g} \times C_{al} \times (-66.29\text{K})$$

$$q_{Al} = C_{al} \times (-331.45\text{gK})$$

3) Since we still can't get the value of C_{al} or q_{Al} , we can turn to the given information for water for help:

$$q_w = 1000.0\text{g} \times 4.184\text{J/gK} \times .71\text{K}$$

$$q_w = 297.06\text{J}$$

4) Following the First Law of Thermodynamics, we can state that:

$$q_{Al} = -q_w$$

$$q_{Al} = -297.06\text{J}$$

5) Put the value of q_{Al} over -331.45gK to get C_{Al}

$$\frac{-297.06 \text{ J}}{-331.45 \text{ g}\cdot\text{K}} = 0.90 \text{ J/gK} \quad (2 \text{ sig. figs. because of } 0.71 \text{ K})$$

General problem solving strategy with calorimetry problems:

- 1) Figure out what the system is. Usually the surroundings will be water (or the whole solution). Usually you will have to figure out the heat change of the surroundings, and then the heat change of the system is opposite that.
- 2) Calculate the heat change using $q = m C \Delta T$. Note that in a system where you have an aqueous solution, the mass (m) will be the mass of the whole solution (solvent + solute), so if it's a solid solute that you're adding to water, you'll have to include the mass of the solute in it.
- 3) If the problem requires you to calculate ΔH as J/mol, then you next have to find out how many moles reacted so you can put it in the denominator, using $\Delta H = q/\text{mol}$
- 4) Make sure the sign of ΔH is correct (+ if the system had an endothermic change, - if the system had an exothermic change)
- 5) Check that the sig figs are correct (ΔT is usually what limits these calculations since in the subtraction to calculate ΔT you often lose sig figs)

Hess' Law

- it states that you can find the enthalpy of a multi-step reaction if you add up all the steps' enthalpy values.
- "products –reactants" equation is special case of Hess' Law, when all the reactions you're using are enthalpies of formation reactions.
- There are good practice problems using Hess' Law in the ACS Guide, pages 50-52, #s13-22.

*There are two chances to attend the Facilitated Study Group (FSG) before Tuesday's Exam 2:

- Friday 3:00-4:00 in Wheatley Rm: 127
- Monday 4:30-5:30 in Wheatley Rm: 127

*There are 2 practice exams (with answer keys) for Exam 2 posted on the Chem. 115 website. They are a really good review of the material that will be on this exam. It would be a good idea to use the practice exam without the multiple choice questions as a study guide or review of everything. Then, with the practice exam that has multiple choice and short answers, try simulating a real test by sitting down and trying to finish it within an hour and fifteen minutes. Good Luck!