Announcements – 9/18/00

- Exam #1

  -info page on website

- Tomorrow's Problem Session

  -put questions in box or email

- Quiz today!

  -but no quiz on Friday this week

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Quantifying Reaction Chemistry

- How many *grams of* O_2 can be produced via the following reaction from *3.0 grams of* $KClO_3$?

$$KClO_3\ (s) \xrightarrow{\Delta} KCl\ (s) + O_2\ (g)\uparrow$$

-First, need a balanced equation:

$$\mathbf{2}KClO_3\ (s) \xrightarrow{\Delta} \mathbf{2}KCl\ (s) + \mathbf{3}O_2\ (g)\uparrow$$

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More QRC

- <u>Next:</u> remember that only MOLES can be used to quantify chemical changes:

g $KClO_3$ ® mol $KClO_3$ ® mol O_2 ® g O_2

$$3.0 \text{ g } KClO_3 \times \frac{1 \text{ mol } KClO_3}{122.548 \text{ g } KClO_3} \times \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} \times \frac{31.998 \text{ g } O_2}{1 \text{ mol } O_2} =$$

= 1.17498 g O_2

= ***<u>1.2 g O_2</u>***

Reaction Reality: Percent Yield

- Previous example gave the *theoretical* yield for the reaction . . . more realistically:

-Suppose the reaction of 3.0 g $KClO_3$ produced 0.55 g O_2; calculate the *percent yield* of the reaction

$$\text{\%-yield} = \frac{\text{Actual (exptl) Yield}}{\text{Theoretical Yield}} \times 100$$

$$= \frac{0.55 \text{ g } O_2}{1.175 \text{ g } O_2} \times 100 = \mathbf{47\%}$$

Limiting Reagent

- We don't always react a *stoichiometric* amount of reactants:

  -How many *g P_2O_5* will be produced by the reaction of *2.00 g P* with *5.00 g O_2*?

  Reaction: $P + O_2 \rightarrow P_2O_5$

  Balance: $\mathbf{4}P + \mathbf{5}O_2 \rightarrow \mathbf{2}P_2O_5$

  Moles: $2.00 \text{ g P} \times \frac{1 \text{ mol P}}{30.974 \text{ g P}} = 0.06457 \text{ mol P}$

  $5.00 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{31.998 \text{ g } O_2} = 0.1563 \text{ mol } O_2$

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Limiting Reagent: Cont'd

- Compare *actual* mol to mol *required*:

$0.06457 \text{ mol P} \times \frac{5 \text{ mol } O_2}{4 \text{ mol P}} = 0.08071 \text{ mol } O_2$

\# mol O_2 needed to react with actual amt of P

So, there will be *O_2 leftover* after all of the P is consumed:

0.1503 mol O_2 - actual

-0.08071 mol O_2 - reacted

0.0756 mol O_2 *unreacted* (excess)

The reaction is limited by the amount of **P**, so it is the ***Limiting Reagent.***

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Limiting Reagent: The Final Straw

- Since **P** is the limiting reagent, we use *its amount* for the final calculation:

g P ® mol P ®mol P_2O_5 ®g P_2O_5

$$2.00\text{ g P} \times \frac{1\text{ mol P}}{30.974\text{ g P}} \times \frac{2\text{ mol }P_2O_5}{4\text{ mol P}} \times \frac{141.943\text{ g }P_2O_5}{1\text{ mol }P_2O_5} =$$

$= 4.58265\text{ g }P_2O_5$

= ***4.58 g P_2O_5***

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Solution Concentrations

- We need to be able to *quantify* amounts of compounds in solutions:

1. **Mass Percent**

$$\textbf{Mass Percent} = \frac{\text{Solute Mass}}{\text{Solution Mass}} \times 100$$

Used, more typically, for very dilute solutions:

$$\textbf{ppm} = \frac{\text{Solute Mass}}{\text{Solution Mass}} \times 10^6 \qquad \textit{Trace}$$

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Concentrations: Moles

- Since reaction chemistry is quantified using moles, these are more useful:

2. **Mole Fraction** (*X*) = $\frac{\text{mol solute}}{\text{Total mol}}$
3. **Molarity (M)** = $\frac{\text{mol solute}}{\text{Liters Solution}}$
4. **Molality (*m*) =** $\frac{\text{mol solute}}{\text{kg solvent}}$

   -temp independent

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