

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 O₂* can be produced via the following reaction from *3.0 grams of KClO₃*?



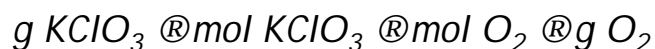
-First, need a balanced equation:



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More QRC

- Next: remember that only MOLES can be used to quantify chemical changes:



$$\begin{aligned} 3.0 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.548 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \times \frac{31.998 \text{ g O}_2}{1 \text{ mol O}_2} &= \\ &= 1.17498 \text{ g O}_2 \\ &= \underline{\underline{1.2 \text{ g O}_2}} \end{aligned}$$

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Reaction Reality: Percent Yield

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

-Suppose the reaction of 3.0 g KClO₃ produced 0.55 g O₂; calculate the *percent yield* of the reaction

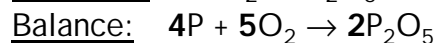
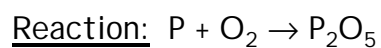
$$\begin{aligned} \text{\% -yield} &= \frac{\text{Actual (exptl) Yield}}{\text{Theoretical Yield}} \times 100 \\ &= \frac{0.55 \text{ g O}_2}{1.175 \text{ g O}_2} \times 100 = \underline{\underline{47\%}} \end{aligned}$$

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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 ?



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:

$$\begin{array}{r} 0.1503 \text{ mol } O_2 \text{ - actual} \\ -0.08071 \text{ mol } O_2 \text{ - reacted} \\ \hline \mathbf{0.0756 \text{ mol } O_2 \text{ unreacted (excess)}} \end{array}$$

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:

$$\begin{aligned} & g P \text{ @ } mol P \text{ @ } mol P_2O_5 \text{ @ } g P_2O_5 \\ 2.00 \text{ g P} & \times \frac{1 \text{ mol P}}{30.974 \text{ g P}} \times \frac{2 \text{ mol } P_2O_5}{4 \text{ mol P}} \times \frac{141.943 \text{ g } P_2O_5}{1 \text{ mol } P_2O_5} = \\ & = 4.58265 \text{ g } P_2O_5 \\ & = \underline{\underline{4.58 \text{ g } P_2O_5}} \end{aligned}$$

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Solution Concentrations

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

1. **Mass Percent**

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

Used, more typically, for very dilute solutions:

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

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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