CHEMISTRY 31
Summer, 2016 - Dixon
Solutions to Additional Problems 1.1 and 1.2

Additional Problem 1.1 (3 points)

The following method is used to capture, oxidize and titrate sulfur dioxide (SO_2) in air in order to measure its concentration. Air is bubbled into 25.0 mL of an 0.0050 M NaOH solution at a rate of 1.00 L (air)/min for 151 min. Excess hydrogen peroxide is present which converts the sulfur dioxide into sulfuric acid through the following reaction:

$$SO_2(g) + H_2O_2(aq) \leftrightarrow H_2SO_4(aq) + H_2O(l)$$

The sulfuric acid is neutralized by the OH^-. The excess NaOH (that not reacted with H_2SO_4), then is titrated with HCl, and required 38.1 mL of 0.00233 M HCl. Determine the concentration of SO_2 in the air in nmol/L.

This problem is somewhat difficult because both OH^- and H_2O_2 are in excess. However, the back titration is with HCl (reacting with OH^-) so we need to consider the moles of OH^-.

$n(OH^-)_{present} = n(OH^-)_{reacted} + n(OH^-)_{excess}$

$n(OH^-)_{reacted}$ *is related to the concentration of SO_2 through the combination of an acid base reaction with the reaction listed above:*

acid-base rxn: $H_2SO_4(aq) + 2OH^-(aq) \leftrightarrow 2H_2O(l) + SO_4^{2-}(aq)$

combined with above reaction:

$SO_2(g) + H_2O_2(aq) + 2OH^-(aq) \leftrightarrow 3H_2O(l) + SO_4^{2-}(aq)$

$n(SO_2) = [n(OH^-)_{reacted}](1 \text{ mol } SO_2/2 \text{ mol } OH^- \text{ reacted})$

$n(OH^-)_{reacted} = n(OH^-)_{present} - n(OH^-)_{excess}$

with $n(OH^-)_{excess} = n(HCl) = (38.1 \text{ mL})(0.00233 \text{ mmol/mL}) = 0.08877 \text{ mmol}$

$n(OH^-)_{reacted} = (25.0 \text{ mL})(0.0050 \text{ mmol/mL}) - 0.08877 \text{ mmol} = 0.125 - 0.08877 = 0.0362 \text{ mmol}$

(note: only 2 sig figs after subtraction)

$n(SO_2) = (0.0362 \text{ mmol})(1 \text{ mol } SO_2/2 \text{ mol } OH^- \text{ reacted}) = 0.0181 \text{ mmol}$

Conc. $= n(SO_2)/V = (0.0181 \text{ mmol})(10^6 \text{ nmol/mmol})/[(1.00 \text{ L/min})(151 \text{ min})] =$ **120 nmol/L**

Additional Problem 1.2 (4 points)

A chemist wishes to weigh out iron for a reaction but her balance is broken. She decides that she can determine the mass of iron by measuring the length and diameter of an iron rod. The density of iron is 7.86 $\pm$ 0.01 g cm^{-3}, the length of the rod is 32.4 $\pm$ 0.4 mm and the diameter of the rod is 4.16 $\pm$ 0.05 mm.

a) Calculate the **mass of iron added.** The volume of a rod equals $\pi \cdot d^2 \cdot l/4$, where d is the rod diameter and l is the rod length.

mass = volume·density $= (\pi \cdot d^2 \cdot l/4) \cdot$ *density*

$= 3.14159(4.16 \text{ mm})^2(32.4 \text{ mm})(7.86 \text{ g cm}^{-3})(1 \text{ cm}^3/1000 \text{ mm}^3)/4 = 3.4613 \text{ g}$

(note: this needs to be corrected to the correct # sig figs)

once b) is done, we know this = ***3.46 g*** *(to hundredths place – same as unc.)*

b) Calculate the **uncertainty in the mass of iron added**.

This is a "mixed operations" problem because it has both exponents and multiplication/division

The first operation is the exponent

$S_{d^{\wedge}2}/d^2 = 2(S_d/d) = 2(0.05/4.16) = 0.0240$

Next is the rest of the problem $S_{mass}/mass = [(S_{d^2}/d^2)^2 + (S_l/l)^2 + (S_{density}/density)^2]^{0.5}$
$S_{mass}/mass = [(0.024)^2 + (0.4/32.4)^2 + (0.01/7.86)^2]^{0.5} = 0.000732^{0.5} = 0.0271$
$S_{mass} = 0.0271 \cdot (3.4613\ g) = 0.0936\ g =$ ***0.09 g***

c) Which measurement contributed most to the overall uncertainty?
The ***diameter measurement*** *contributed the most to the overall uncertainty because* S_{d^2}/d^2 *was greater than* S_l/l *or* $S_{density}/density$