University of Waterloo
DEPARTMENT OF CHEMISTRY
CHEM 120 - Term Test #1
Wednesday, October 17, 2012

This is test version 001.
Fill in ovals 001 for the Card Number (or Test Master) on your computer answer card.

Name (Print clearly): ______________________________________
First name Surname

Student Number: ____________________ Signature: ______________________________

Instructors: R. Marta (sec. 001, 002); C. Bissonnette (sec. 003, 005); D. Gilbert (sec. 004)

Aids Allowed: Electronic Calculator

Time Allowed: 50 minutes

Instructions:

1. Print your name and other information in the top right corner on the front of the computer card.
2. Fill in your ID number, Section Number and your Card Number/Test Master in the designated fields.
3. **Use a pencil to fill in the ovals on the computer card.** Make sure any erasures are complete and clean. If you need a pencil, raise your hand and ask for one.
4. This test booklet has 16 questions and 8 pages (including this Cover Page and the Data Sheet). Check now to ensure that no pages are missing. **If pages are missing, raise your hand and ask for a new test booklet.**
5. There is only one correct answer for each question. **Circle A, B, C, D or E on this booklet**. When you have finished, **transfer your answers to the computer card** by filling in the appropriate oval (A,B,C,D or E) in the column for that question on the card.
6. Attempt all questions. No marks are deducted for wrong answers. Do all rough work on this test booklet. Scrap paper is not permitted.
7. At the end of the test, turn in this test booklet, the data sheet, and your computer card, and then leave immediately. **Do not put computer cards inside the test booklet!**

Note: The last page is a DATA SHEET. You may tear it off. Return it with your test booklet and computer card

1. (0.5 marks) Balance the equation below using the **smallest whole number coefficients**. What is the coefficient of $K_2Se(aq)$ in the balanced equation?

$$Sb(NO_3)_3(aq) \quad + \quad K_2Se(aq) \quad \rightarrow \quad Sb_2Se_3(s) \quad + \quad KNO_3(aq)$$

A. 6
B. 2
C. 1
D. 4
E. 3*

Beginning from the left (reactant) side of the equation:

$$Sb(NO_3)_3(aq) + K_2Se(aq) \rightarrow Sb_2Se_3(s) + KNO_3(aq)$$

There is one Sb in $Sb(NO_3)_3$ and are 2 Sb's in Sb_2Se_3. Therefore,

$$\mathbf{2}\ Sb(NO_3)_3(aq) + K_2Se(aq) \rightarrow Sb_2Se_3(s) + KNO_3(aq)$$

There are 6 NO_3's in 2 $Sb(NO_3)_3$'s and one NO_3 in KNO_3. Therefore,

$$2\ Sb(NO_3)_3(aq) + K_2Se(aq) \rightarrow Sb_2Se_3(s) + \mathbf{6}\ KNO_3(aq)$$

There are 6 K's in 6 KNO_3's and 2 K's in K_2Se. Therefore,

$$2\ Sb(NO_3)_3(aq) + \mathbf{3}\ K_2Se(aq) \rightarrow Sb_2Se_3(s) + 6\ KNO_3(aq)$$

There are 3 Se's in 3 K_2Se's and 3 Se's in Sb_2Se_3; the equation is balanced.

2. (0.5 marks) A 4.169 g sample of a pure element contains 3.351×10^{22} atoms. What is the element?

A. K
B. Mo
C. As*
D. Os
E. Pd

The molar mass ($\mathcal{M}$) of an element can be used to identify it.

$\mathcal{M} = \dfrac{m}{n}$ where *m* and *n* are the mass and moles of the element, respectively.

The mass of the element is 4.169 g, and the moles can be obtained from the number of atoms (*N*) using Avogadro's number (N_A).

$$n = \frac{N}{N_A} = \frac{3.351 \times 10^{22} \text{ atoms}}{6.022 \times 10^{23} \text{ atoms mol}^{-1}} = 0.05565 \text{ mol}$$

Therefore,

$$\mathcal{M} = \frac{m}{n} = \frac{4.169 \text{ g}}{0.05565 \text{ mol}} = 74.92 \text{ g mol}^{-1}$$

The molar mass matches best for arsenic, As (from the data sheet).

3. (0.5 marks) What mass of Cr is found in 19.6 g of $K_2Cr_2O_7$? Choose the closest answer.

A. 6.93 g*
B. 8.11 g
C. 13.9 g
D. 5.71 g
E. 3.46 g

We can use the **mole process** to find the mass of Cr in $K_2Cr_2O_7$.

A balanced "atomization" equation (hypothetical) can be written as an aid to recognize that there are 2 moles of Cr in every 1 mole of $K_2Cr_2O_7$.

$$K_2Cr_2O_7 \rightarrow 2\text{ K} + 2\text{ Cr} + 7\text{ O}$$

Use the mole process to find the moles of Cr from the moles of $K_2Cr_2O_7$.

To moles:

$$n_{K_2Cr_2O_7} = \frac{m_{K_2Cr_2O_7}}{\mathcal{M}_{K_2Cr_2O_7}} = \frac{19.6\text{ g}}{294.2\text{ g mol}^{-1}} = 0.06662\text{ mol}$$

Between moles:

There are 2 moles of Cr for every 1 mole of $K_2Cr_2O_7$.

$$n_{Cr} = n_{K_2Cr_2O_7}\left(\frac{2}{1}\right) = 0.06662\text{ mol}\left(\frac{2}{1}\right) = 0.1332\text{ mol}$$

From moles:

$$m_{Cr} = n_{Cr}\mathcal{M}_{Cr} = 0.1332\text{ mol}(52.00\text{ g mol}^{-1}) = 6.93\text{ g}$$

4. (0.5 marks) Suppose that a solid mixture containing 0.01 mol each of $CaCl_2$ and $Fe(NO_3)_2$ is added to 1.0 L of water. Which of the following best describes what happens?

A. $Fe(NO_3)_2$ dissolves but none of the $CaCl_2$ dissolves.
B. $CaCl_2$ dissolves but none of the $Fe(NO_3)_2$ dissolves.
C. A precipitate of $FeCl_2$ is formed.
D. A precipitate of $Ca(NO_3)_2$ is formed.
E. Both $CaCl_2$ and $Fe(NO_3)_2$ dissolve.*

This question addresses the solubility rules of ionic species that you are responsible for memorizing.

Decide if the ionic salts dissolve by considering if they are soluble (dissolve and become aqueous) or insoluble (form a solid precipitate).

1. Ionic salts containing nitrate (NO_3^-) are soluble, therefore **$Fe(NO_3)_2$ will dissolve.**
2. Ionic salts containing anions (negative ions) of halides (F^-, Cl^-, Br^- or I^-) are soluble, except those formed with cations (positive ions) of the metals Hg^+/Hg^{2+}, Pb^{2+} and Ag^+. **$CaCl_2$ will dissolve.**

5. (0.5 marks) What volume of 0.117 mol L^{-1} HNO_3(aq) is needed to prepare by dilution exactly 525 mL of 0.0500 mol L^{-1} HNO_3(aq)?

A. 4.46 mL
B. 224 mL*
C. 35.2 mL
D. 61.4 mL
E. 352 mL

The moles of HNO_3(aq) required are

$$n_{\mathrm{HNO_3},required} = C_{final}\mathrm{V}_{\mathrm{HNO_3},\,final} = 0.0500 \text{ mol L}^{-1}(0.525 \text{ L}) = 0.02625 \text{ mol}$$

Therefore the volume required from the initial solution is,

$$\mathrm{V}_{\mathrm{HNO_3},initial} = \frac{n_{\mathrm{HNO_3},required}}{C_{initial}} = \frac{0.02625 \text{ mol}}{0.117 \text{ mol L}^{-1}} = 0.2244 \text{ L} = 224 \text{ mL}$$

6. (0.5 marks) Which of the following statements best describes the F^{-} ion in the following reaction? Assume the reaction occurs in aqueous solution.

$$F^{-} + HNO_2 \rightarrow HF + NO_2^{-}$$

A. The F^{-} ion causes N in HNO_2 to be oxidized.
B. The F^{-} ion causes N in HNO_2 to be reduced.
C. The F^{-} ion acts as a Bronsted-Lowry acid.
D. The F^{-} ion acts as a Bronsted-Lowry base.*
E. The F^{-} ion acts as a precipitating agent.

A precipitation reaction involves the formation of an insoluble ionic solid (based on the solubility rules) from aqueous reactants.

The reaction shown above does not contain insoluble ionic solids as any of its products. This is not a precipitation reaction.

An acid-base reaction must involve the transfer of one or more protons (H^{+}) from one reactant species to another.

The reaction shown above involves F^{-} removing (accepting) a proton (H^{+}) from HNO_2. F^{-} is acting as a Bronsted-Lowry base. **The correct answer is choice D.**

An oxidation-reduction (redox) reaction must involve the change in oxidation states of one or more of the elements contained in the reactant species.

If oxidation states are assigned to the elements in the reaction shown above, you will find that all elements maintain the same oxidation number between the reactant and product species. It is not a redox reaction.

7. (0.5 marks) If 1.20 moles of an ideal gas occupy 20.5 L at 25.0 ^{o}C, what is the gas pressure?

A. 145 kPa*
B. 12.2 kPa
C. 183 kPa
D. 22.5 kPa
E. 275 kPa

For an ideal gas:

$$PV = nRT, \quad P = \frac{nRT}{V}$$

$$P = \frac{1.20\ \text{mol}(8.3145\ \text{kPa L mol}^{-1}\ \text{K}^{-1})(298.15\ \text{K})}{20.5\ \text{L}} = 145\ \text{kPa}$$

8. (0.5 marks) Both neon (Ne) and argon (Ar) behave as ideal gases at 298 K and 100 kPa. For these conditions of temperature and pressure, a ~~2.00~~ 16.3 g sample of Ne(g) occupies a volume of 20.0 L. For the same conditions, a ~~2.00~~ 16.3 g sample of Ar(g) occupies a volume

A. equal to 20.0 L
B. greater than 20.0 L
C. less than 20.0 L*

The volume (*V*) of an ideal gas (*X*), at constant *T* and *P*, is found using the ideal gas equation.

$$PV_X = n_X RT$$

$$V_X = \frac{n_X RT}{P}$$

The two ideal gases in each sample have **equal masses (*m*)**, and since

$$n_X = \frac{m}{\mathcal{M}_X}, \text{ then}$$

$$V_X = \frac{mRT}{\mathcal{M}_X P} = \frac{k}{\mathcal{M}_X}, \quad k = \frac{mRT}{P} \text{ (constant)}$$

If $V_X = \frac{k}{\mathcal{M}_X}$, then the volume ($V_X$) occupied by a gaseous species is inversely proportional to its molar mass ($\mathcal{M}_X$).

Therefore, the species with a **greater molar mass** will occupy a **smaller volume** at constant *m*, *T* and *P*. **Ar has a greater molar mass than Ne, so it occupies a volume less than 20.0 L; at constant *m*, *T* and *P*.**

NOTE: Although it was not necessary to perform calculations in this problem, the mass of each gas was intended to be 16.3 g. Because of this error, we have awarded any choice of answer as correct!

9. (1 mark) What is the minimum mass of oxygen required to completely burn 13.1 g of $C_7H_6O_2$? The **unbalanced** chemical equation is shown below.

$$C_7H_6O_2(s) \quad + \quad O_2(g) \quad \rightarrow \quad CO_2(g) \quad + \quad H_2O(g)$$

Molar masses: (in g mol^{-1}) $C_7H_6O_2$, 122.12 O_2, 32.00 CO_2, 44.01 H_2O, 18.02

A. 25.7 g*
B. 34.0 g
C. 0.458 g
D. 14.5 g
E. 3.43 g

We can use the **mole process** to find the minimum mass of O_2 required.

First, balance the unbalanced chemical equation.

$$C_7H_6O_2(s) \quad + \quad O_2(g) \quad \rightarrow \quad CO_2(g) \quad + \quad H_2O(g)$$

There are 7 C's on the left side (LS) and one C on the right side (RS).

$$C_7H_6O_2(s) \quad + \quad O_2(g) \quad \rightarrow \quad \mathbf{7}\ CO_2(g) \quad + \quad H_2O(g)$$

There are 6 H's on the LS and 2 H's on the RS.

$$C_7H_6O_2(s) \quad + \quad O_2(g) \quad \rightarrow \quad 7\ CO_2(g) \quad + \quad \mathbf{3}\ H_2O(g)$$

There are 4 O's on the LS and 17 O's on the RS. The equation is balanced.

$$C_7H_6O_2(s) \quad + \quad \frac{\mathbf{15}}{\mathbf{2}}\ O_2(g) \quad \rightarrow \quad 7\ CO_2(g) \quad + \quad 3\ H_2O(g)$$

Use the mole process to find the moles of O_2 from the moles of $C_7H_6O_2$.

To moles:

$$n_{C_7H_6O_2} = \frac{m_{C_7H_6O_2}}{\mathcal{M}_{C_7H_6O_2}} = \frac{13.1\text{ g}}{122.12\text{ g mol}^{-1}} = 0.1073\text{ mol}$$

Between moles:

There are $\frac{15}{2}$ moles of O_2 reacting for every one mole of $C_7H_6O_2$.

$$n_{O_2} = n_{C_7H_6O_2}\left(\frac{15}{2}\right) = 0.1073\text{ mol}\left(\frac{15}{2}\right) = 0.8045\text{ mol}$$

From moles:

$$m_{O_2} = n_{O_2}\mathcal{M}_{O_2} = 0.8045\text{ mol}(32.00\text{ g mol}^{-1}) = 25.7\text{ g}$$

10. (1 mark) What is the maximum mass of $Ca(CN)_2$ that can be obtained from 1.56 g of HCN and 2.58 g of $Ca(OH)_2$? The balanced chemical equation is shown below. (Choose the closest answer.)

$$2\ HCN(aq) + Ca(OH)_2(aq) \rightarrow Ca(CN)_2(aq) + 2\ H_2O(l)$$

Molar masses: (in g mol^{-1})	HCN, 27.03	$Ca(OH)_2$, 74.10	$Ca(CN)_2$, 92.12	H_2O, 18.02

A. 5.31 g
B. 2.20 g
C. 2.66 g*
D. 1.37 g
E. 3.21 g

There are fixed amounts (in grams) of reagents given in the question; this is potentially a limiting reagent problem.

We can use the **mole process** to find the amount of $Ca(CN)_2$ produced by first assuming that all of the HCN reacts, then by assuming that all of the $Ca(OH)_2$ reacts. The reagent that yields the smaller quantity of $Ca(CN)_2$ is the limiting reagent. The amount of the limiting reagent dictates the amount of product obtained.

To moles:

$$n_{HCN} = \frac{m_{HCN}}{\mathcal{M}_{HCN}} = \frac{1.56\ g}{27.03\ g\ mol^{-1}} = 0.05771\ mol$$

Between moles:

There are two moles of HCN reacting to form one mole of $Ca(CN)_2$.

$$n_{Ca(CN)_2} = n_{HCN}\left(\frac{1}{2}\right) = 0.05771\ mol\left(\frac{1}{2}\right) = 0.02886\ mol$$

To moles:

$$n_{Ca(OH)_2} = \frac{m_{Ca(OH)_2}}{\mathcal{M}_{Ca(OH)_2}} = \frac{2.58\ g}{74.10\ g\ mol^{-1}} = 0.03482\ mol$$

Between moles:

There are two moles of HCN reacting to form one mole of $Ca(CN)_2$.

$$n_{Ca(CN)_2} = n_{Ca(OH)_2}\left(\frac{1}{1}\right) = 0.03482\ mol$$

The limiting reagent is HCN because it produces less $Ca(CN)_2$, therefore,

From moles:

$$m_{Ca(CN)_2} = n_{Ca(CN)_2}\mathcal{M}_{Ca(CN)_2} = 0.02886\ mol(92.12\ g\ mol^{-1}) = 2.66\ g$$

11. (1 mark) A 4.23 g sample of a mixture of $BaCl_2$ and $Ba(NO_3)_2$ is treated with excess $AgNO_3(aq)$. The precipitate is filtered off and then dried. If the dry precipitate weighs 5.27 g, what is the percentage by mass of $BaCl_2$ in the original mixture?

Molar masses: $BaCl_2$, 208.2 $AgNO_3$, 169.91 AgCl, 143.35 $Ba(NO_3)_2$, 261.32
(in g mol^{-1})

A. 63.3%
B. 90.5%*
C. 21.5%
D. 35.2%
E. 45.6%

There is no balanced chemical equation given in the problem. We must decide what reaction(s) is occurring, then write the balanced chemical equation.

First focus on $BaCl_2$ reacting with $AgNO_3$. AgCl is formed as a solid precipitate (based on our solubility rules).

$$BaCl_2(aq) + AgNO_3(aq) \rightarrow AgCl(s) + Ba(NO_3)_2(aq)$$

Balance the equation.

There are 2 Cl's on the LS and one 1 Cl on the RS.

$$BaCl_2(aq) + AgNO_3(aq) \rightarrow \mathbf{2}\ AgCl(s) + Ba(NO_3)_2(aq)$$

There are 2 Ag's on the RS and one 1 Ag on the LS. The equation is balanced.

$$BaCl_2(aq) + \mathbf{2}\ AgNO_3(aq) \rightarrow 2\ AgCl(s) + Ba(NO_3)_2(aq)$$

You can attempt to write a balanced chemical equation for a potential reaction between $Ba(NO_3)_2$ and $AgNO_3$, but should find that no reaction will occur!

We can use the **mole process** to find the mass of $BaCl_2$ in the sample.

Use the mole process to find the moles of $BaCl_2$ from the moles of AgCl.

To moles:

$$n_{\text{AgCl}} = \frac{m_{\text{AgCl}}}{\mathcal{M}_{\text{AgCl}}} = \frac{5.27\text{ g}}{143.35\text{ g mol}^{-1}} = 0.03676\text{ mol}$$

Between moles:

There is one mole of $BaCl_2$ consumed for every two moles of AgCl formed.

$$n_{\text{BaCl}_2} = n_{\text{AgCl}}\left(\frac{1}{2}\right) = 0.03676\text{ mol}\left(\frac{1}{2}\right) = 0.01838\text{ mol}$$

From moles:

$$m_{\text{BaCl}_2} = n_{\text{BaCl}_2}\mathcal{M}_{\text{BaCl}_2} = 0.01838\text{ mol}(208.2\text{ g mol}^{-1}) = 3.827\text{ g}$$

% of $BaCl_2$ in sample:

$$\%_{\text{BaCl}_2} = \frac{m_{\text{BaCl}_2}}{m_{\text{sample}}} \times 100\% = \frac{3.827\text{ g}}{4.23\text{ g}} \times 100\% = 90.5\%$$

12. (1 mark) The concentration of Fe^{2+} in an aqueous solution can be determined by titration with $K_2Cr_2O_7$(aq) in the presence of excess acid. Titration of a 25.00 mL sample of a solution of Fe^{2+} requires 28.72 mL of 0.05051 mol L^{-1} $K_2Cr_2O_7$(aq) to reach the endpoint. What is the concentration of Fe^{2+} in the original solution? The net ionic equation for the reaction is given below. (Choose the closest answer.)

$$6\ Fe^{2+}(aq) + 14\ H^{+}(aq) + Cr_2O_7^{2-}(aq) \rightarrow 6\ Fe^{3+}(aq) + 2\ Cr^{3+}(aq) + 7\ H_2O(l)$$

A. 0.05803 mol L^{-1}

B. 0.3482 mol L^{-1}*

C. 0.009671 mol L^{-1}

D. 0.1741 mol L^{-1}

E. 0.02901 mol L^{-1}

The $Cr_2O_7^{2-}$ can only come directly from $K_2Cr_2O_7$.

$$K_2Cr_2O_7(aq) \xrightarrow{100\%} 2\ K^{+}(aq) + Cr_2O_7^{2-}(aq)$$

For every one $K_2Cr_2O_7$ that dissociates there is one $Cr_2O_7^{2-}$ produced, so

$$n_{K_2Cr_2O_7} = n_{Cr_2O_7^{2-}} \text{ or } C_{K_2Cr_2O_7} = C_{Cr_2O_7^{2-}}$$

We can use the **mole process** to find the concentration of Fe^{2+} in the sample.

Use the mole process to find the moles of Fe^{2+} from the moles of $Cr_2O_7^{2-}$.

To moles:

$$n_{Cr_2O_7^{2-}} = C_{Cr_2O_7^{2-}} V_{Cr_2O_7^{2-}} = 0.05051 \text{ mol L}^{-1}(0.02872 \text{ L}) = 1.451\times10^{-3} \text{ mol}$$

Between moles:

There are 6 moles of Fe^{2+} consumed for every one mole of $Cr_2O_7^{2-}$ consumed.

$$n_{Fe^{2+}} = n_{Cr_2O_7^{2-}}\left(\frac{6}{1}\right) = 1.451\times10^{-3} \text{ mol}(6) = 8.704\times10^{-3} \text{ mol}$$

From moles:

$$C_{Fe^{2+}} = \frac{n_{Fe^{2+}}}{V_{Fe^{2+}}} = \frac{8.704\times10^{-3} \text{ mol}}{0.02500 \text{ L}} = 0.03482 \text{ mol L}^{-1}$$

13. (1 mark) In aqueous solution, ammonium nitrate, NH_4NO_3, and sodium hydroxide, NaOH, react to give $NaNO_3$, NH_3, and water. What is the **<u>net ionic equation</u>** for this reaction?

A. $Na^+(aq) + NO_3^-(aq) \rightarrow NaNO_3(s)$

B. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

C. $NH_4NO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l) + NH_3(aq)$

D. $NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(l)$*

E. $H^+(aq) + NO_3^-(aq) \rightarrow HNO_3(aq)$

A balanced chemical equation is not given in the problem. Write a balanced chemical equation based on what is described by the question.

$NH_4NO_3 + NaOH \rightarrow NaNO_3 + NH_3 + H_2O$

The equation is balanced.

Next, decide if any of the species will dissolve and/or dissociate (using solubility rules, strong/weak electrolytes, etc...).

NH_4NO_3 is a soluble ionic compound, yielding NH_4^+ and NO_3^- ions in solution.

NaOH is a soluble ionic compound, yielding Na^+ and OH^- ions in solution.

$NaNO_3$ is a soluble ionic compound, yielding Na^+ and NO_3^- ions in solution.

Ammonia, NH_3 is a molecular compound that is not in our lists of strong acids or bases, so we can only assume that is a weak or non-electrolyte (it is a weak base and thus, weak electrolyte). Therefore NH_3 is best represented as $NH_3(aq)$.

Liquid water, $H_2O(l)$, does not dissociate appreciably. Represent H_2O as $H_2O(l)$.

$NH_4NO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + NH_3(aq) + H_2O(l)$

Separate species that dissociate fully into their constituent ions.
$NH_4^+(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + NH_3(aq) + H_2O(l)$

Remove spectator ions.
$NH_4^+(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + NH_3(aq) + H_2O(l)$

Done.
$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(l)$ (an acid-base neutralization reaction)

14. (1 mark) Consider the unbalanced redox reaction shown below.

$$IO_3^-(aq) \;+\; HSO_3^-(aq) \;\rightarrow\; I^-(aq) \;+\; SO_4^{2-}(aq) \qquad \text{(in } \textit{acidic} \text{ solution)}$$

When the equation is balanced with the **smallest whole number coefficients**, what is the coefficient of SO_4^{2-} ?

A. 1

B. 2

C. 3*

D. 4

E. 5

Assign oxidation states to the elements in the reaction.

The oxidation state of O in a compound is −2 (not in O_2) unless it is contained in hydrogen peroxide (H_2O_2) or is attached to fluorine (e.g. F_2O, F_2O_2, etc...).

The oxidation state of H in a compound is +1 (not in H_2) unless it is contained in a metal hydride (e.g. NaH, LiH, CaH_2, etc...), where H has an oxidation state of −1.

$$\sum \text{oxidation \#'s} = \text{net charge}$$

Let *x* be the oxidation # of **I in IO_3^-**.

$x + 3(-2) = -1$

$x = +5$

Let *x* be the oxidation # of **I in I^-**.

$x = -1$

Let *x* be the oxidation # of **S in HSO_3^-**.

$1 + x + 3(-2) = -1$

$x = +4$

Let *x* be the oxidation # of **S in SO_4^{2-}**.

$x + 4(-2) = -2$

$x = +6$

Determine what species are being oxidized and reduced.

Since the oxidation # of I in IO_3^-, changes from +5 to −1 in I^-, it has become more negative, thus I has **gained** (GER or RIG) 6 electrons ($6e^-$) and is **reduced**.
Since the oxidation # of S in HSO_3^-, changes from +4 to a +6 in SO_4^{2-}, it has become more positive, thus S has **lost** (LEO or OIL) 2 electrons ($2e^-$) and is **oxidized**.

Write down the half reactions for the reduction and oxidation processes.

red: $IO_3^-(aq) \rightarrow I^-(aq)$
ox: $HSO_3^-(aq) \rightarrow SO_4^{2-}(aq)$

Balance the elements being oxidized and reduced where applicable.

The elements being oxidized and reduced are already balanced.

Add in the electrons being transferred (be careful to account for stoichiometry where applicable).

The I in IO_3^- **gains** $6e^-$ upon forming I^-
The S in HSO_3^- **loses** $2e^-$ upon forming SO_4^{2-}

red: $IO_3^-(aq) + 6\,e^- \rightarrow I^-(aq)$
ox: $HSO_3^-(aq) \rightarrow SO_4^{2-}(aq) + 2\,e^-$

14. continued

Combine the half reactions in such a way that the electrons cancel.

Leave the reduction half reaction as is, then add it to 3 times the oxidation half reaction.

ox: $[HSO_3^-(aq) \rightarrow SO_4^{2-}(aq) + 2\,e^-] \times 3$
ox: $3\,HSO_3^-(aq) \rightarrow 3\,SO_4^{2-}(aq) + 6\,e^-$

Therefore,

red: $IO_3^-(aq) + 6\,e^- \rightarrow I^-(aq)$
ox: $\underline{3\,HSO_3^-(aq) \rightarrow 3\,SO_4^{2-}(aq) + 6\,e^-}$

gives,

$$IO_3^-(aq) + 3\,HSO_3^-(aq) \rightarrow I^-(aq) + \mathbf{3}\,SO_4^{2-}(aq)$$

The stoichiometric coefficient in front of SO_4^{2-} will not change by balancing the equation with H^+(aq), then H_2O(l). Done.

15. (1 mark) A volatile liquid produces 0.339 g of vapour in a 0.225 L flask at 100.0 °C and 100.0 kPa. What is the molar mass of the liquid?

A. 61.1 g mol^{-1}
B. 123 g mol^{-1}
C. 42.0 g mol^{-1}
D. 309 g mol^{-1}
E. 46.7 g mol^{-1}*

For an ideal gas:

$$PV = nRT, \quad n = \frac{m}{\mathcal{M}}$$

where m and $\mathcal{M}$ are the mass and molar mass of the vaporized liquid, respectively.

$$n = \frac{PV}{RT} = \frac{100.0\text{ kPa}(0.225\text{ L})}{8.3145\text{ kPa L K}^{-1}\text{ mol}^{-1}(373.15\text{ K})} = 7.252\times10^{-3}\text{ mol}$$

$$\mathcal{M} = \frac{m}{n} = \frac{0.339\text{ g}}{7.252\times10^{-3}\text{ mol}} = 46.7\text{ g mol}^{-1}$$

16. (1 mark) In an experiment, a 0.619 g sample of Al(s) was completely dissolved in an excess of HCl(aq). The H_2 gas produced by the reaction below is then dried to remove all the water vapour. The “dry” H_2(g) occupied a volume of 722 mL at 23.0 ^{o}C and 99.3 kPa. What was the percent yield for this experiment? (Choose the closest answer.)

$$2\ \text{Al(s)}\ +\ 6\ \text{HCl(aq)}\ \rightarrow\ 2\ \text{AlCl}_3\text{(aq)}\ +\ 3\ \text{H}_2\text{(g)}$$

A. 38%
B. 56%
C. 67%
D. 85%*
E. 92%

The percent yield is given by:

$$\%\ \text{yield} = \frac{m_{\text{actual}}}{m_{\text{theory}}} \times 100\%$$

We can use the **mole process** to find the **theoretical mass** (100% yield) of H_2 produced.

Use the mole process to find the moles of H_2 from the moles of Al.

To moles:

$$n_{\text{Al}} = \frac{m_{\text{Al}}}{\mathcal{M}_{\text{Al}}} = \frac{0.619\ \text{g}}{26.98\ \text{g mol}^{-1}} = 0.02294\ \text{mol}$$

Between moles:

There are 2 moles of Al consumed for every 3 moles of H_2 formed.

$$n_{\text{H}_2} = n_{\text{Al}}\left(\frac{3}{2}\right) = 0.02294\ \text{mol}\left(\frac{3}{2}\right) = 0.03441\ \text{mol}$$

From moles:

$$m_{\text{H}_2} = n_{\text{H}_2}\mathcal{M}_{\text{H}_2} = 0.03441\ \text{mol}(2.016\ \text{g mol}^{-1}) = 0.06937\ \text{g}$$

We can use the **experimental conditions** described in the question and the ideal gas equation to find the **actual mass** of H_2 produced.

To moles:

$$PV = nRT,\quad n = \frac{PV}{RT} = \frac{99.3\ \text{kPa}(0.722\ \text{L})}{8.3145\ \text{kPa L K}^{-1}\ \text{mol}^{-1}(296.15\ \text{K})} = 0.02912\ \text{mol}$$

From moles:

$$m_{\text{H}_2} = n_{\text{H}_2}\mathcal{M}_{\text{H}_2} = 0.02912\ \text{mol}(2.016\ \text{g mol}^{-1}) = 0.05870\ \text{g}$$

$$\%\ \text{yield} = \frac{m_{\text{actual}}}{m_{\text{theory}}} \times 100\% = \frac{0.05870\ \text{g}}{0.06937\ \text{g}} \times 100\% = 85.0\%$$

Another (more elegant) approach is to calculate the volume of H_2 expected from 0.619 g of Al and find the % yield as follows:

$$\%\text{ yield} = \frac{V_{\text{actual}}}{V_{\text{theory}}} \times 100\%$$

We can use the **mole process** to find the **theoretical volume** (100% yield) of H_2 produced. The **actual volume** of H_2 produced is 722 mL.

Use the mole process to find the moles of H_2 from the moles of Al.

To moles:

$$n_{\text{Al}} = \frac{m_{\text{Al}}}{\mathcal{M}_{\text{Al}}} = \frac{0.619\text{ g}}{26.98\text{ g mol}^{-1}} = 0.02294\text{ mol}$$

Between moles:

There are 2 moles of Al consumed for every 3 moles of H_2 formed.

$$n_{\text{H}_2} = n_{\text{Al}}\left(\frac{3}{2}\right) = 0.02294\text{ mol}\left(\frac{3}{2}\right) = 0.03441\text{ mol}$$

From moles:

$$PV = nRT, \quad V = \frac{nRT}{P} = \frac{0.03441\text{ mol}(8.3145\text{ kPa L K}^{-1}\text{ mol}^{-1})(296.15\text{ K})}{99.3\text{ kPa}} = 0.8533\text{ L} = 853.3\text{ mL}$$

$$\%\text{ yield} = \frac{722\text{ mL}}{853.3\text{ mL}} \times 100\% = 84.6\%$$