THE COPPERBELT UNIVERSITY

DEPARTMENT OF CHEMISTRY

CH110: TUTORIAL SHEET 2

REACTIONS IN AQUEOUS SOLUTIONS

Term I (2023-2024)

1. The compounds K_2CO_3, Na_2CO_3, KCl and NaCl are soluble in water, but $CaCO_3$ is not. Given the following molecular equation:

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaCl(aq)$$

   (a) Write *molecular equation* for the reaction between solutions of $CaCl_2$ and K_2CO_3
   (b) Write the *ionic equation* for the reaction between solutions of $CaCl_2$ and K_2CO_3
   (c) What is meant by the term *spectator ion*
   (d) Write down the spectator ion(s) for this reaction if any
   (e) Write the *net ionic equation* for the reaction between solutions of $CaCl_2$ and K_2CO_3

2. What volume of 16 *M* sulfuric acid must be used to prepare 1.5 L of a 0.10-*M* H_2SO_4 solution?

3. Assign oxidation numbers to all the elements in the following compounds and ion: (a) Na_2O, (b) HNO_2, (c) $Cr_2O_7^{-2}$

4. Define molarity. Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl in enough water to make 26.8 mL of solution.

5. Glycine (H_2NCH_2COOH) is the simplest amino acid. What is the molarity of an aqueous solution that contains 0.715 mol of glycine in 495 mL?

6. Specialized cells in the stomach release HCl to aid digestion. If they release too much, the excess can be neutralized with an antacid to avoid discomfort. A common antacid contains magnesium hydroxide, $Mg(OH)_2$, which reacts with the acid to form water and magnesium chloride solution. As a government chemist testing commercial antacids, you use 0.10 *M* HCl to simulate the acid concentration in the

stomach. How many liters of "stomach acid" react with a tablet containing 0.10 g of $Mg(OH)_2$?

7. Predict whether a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, total ionic, and net ionic equations, and identify the spectator ions.
   (a) Potassium fluoride(*aq*) + strontium nitrate(*aq*) →
   (b) Ammonium perchlorate(*aq*) + sodium bromide(*aq*) →

8. You perform an acid-base titration to standardize an HCl solution by placing 50.00 mL of HCl in a flask with a few drops of indicator solution. You put 0.1524 *M* NaOH into the buret, and the initial reading is 0.55 mL. At the end point, the buret reading is 33.87 mL. What is the concentration of the HCl solution?

9. Identify the oxidizing agent and reducing agent in each of the following:
   (a) $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$
   (b) $PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$
   (c) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

10. Permanganate ion is a strong oxidizing agent, and its deep purple color makes it useful as an indicator in redox titrations. It reacts in basic solution with the oxalate ion to form carbonate ion and solid manganese dioxide. Balance the skeleton ionic equation for the reaction between $NaMnO_4$ and $Na_2C_2O_4$ in basic solution:
$$MnO_4^-(aq) + C_2O_4^{2-}(aq) \rightarrow MnO_2(s) + CO_3^{2-}(aq)$$

11. Balance the redox reaction between dichromate ion and iodide ion to form chromium(III) ion and solid iodine, which occurs in acidic solution.
$$Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$$

Solutions

Solution 1:

(a) Write molecular equation for the reaction between solutions of $CaCl_2$ and K_2CO_3

$$CaCl_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2KCl(aq)$$

(b) Write the ionic equation for the reaction between solutions of $CaCl_2$ andK_2CO_3

$$Ca^{2+}(aq) + 2Cl^-(aq) + 2K^+(aq) + CO_3^-(aq) \rightarrow CaCO_3(s) + 2Cl^-(aq) + 2K^+(aq)$$

(c) What is meant by the term spectator ion

A **spectator ion** is an **ion** that exists as a reactant and a product in a chemical equation

(d) Write down the spectator ion(s)for this reaction if any

$$2Cl^-(aq) + 2K^+(aq) \rightarrow 2Cl^-(aq) + 2K^+(aq)$$

(e) Write the net ionic equation for the reaction between solutions of $CaCl_2$ and K_2CO_3

$$Ca^{2+}(aq) + CO_3^-(aq) \rightarrow CaCO_3(s)$$

Solution 2:

Solution

Where are we going?

To find the volume of H_2SO_4 required to prepare the solution

What do we know?

- 1.5 L of 0.10 *M* H_2SO_4 is required
- We have 16 *M* H_2SO_4

What information do we need to find the volume of H_2SO_4?

- Moles of H_2SO_4 in the required solution

How do we get there?

What are the moles of H_2SO_4 required?

$$M \times V = \text{mol}$$

$$1.5\ \cancel{\text{L solution}} \times \frac{0.10 \text{ mol } H_2SO_4}{\cancel{\text{L solution}}} = 0.15 \text{ mol } H_2SO_4$$

What volume of 16 M H_2SO_4 contains 0.15 mole of H_2SO_4?

$$V \times \frac{16 \text{ mol } H_2SO_4}{\text{L solution}} = 0.15 \text{ mol } H_2SO_4$$

Solving for *V* gives

$$V = \frac{0.15\ \cancel{\text{mol } H_2SO_4}}{\dfrac{16\ \cancel{\text{mol } H_2SO_4}}{1 \text{ L solution}}} = 9.4 \times 10^{-3} \text{ L or } 9.4 \text{ mL solution}$$

Solution 3:

Solution

(a) By rule 2, we see that sodium has an oxidation number of +1 (Na^+) and oxygen's oxidation number is −2 (O^{2-}).

(b) This is the formula for nitrous acid, which yields a H^+ ion and a NO_2^- ion in solution. From rule 4, we see that H has an oxidation number of +1. Thus, the other group (the nitrite ion) must have a net oxidation number of −1. Oxygen has an oxidation number of −2, and if we use x to represent the oxidation number of nitrogen, then the nitrite ion can be written as

$$[N^{(x)}O_2^{(2-)}]^-$$

so that $$x + 2(-2) = -1$$

or $$x = +3$$

(c) From rule 6, we see that the sum of the oxidation numbers in the dichromate ion $Cr_2O_7^{2-}$ must be −2. We know that the oxdiation number of O is −2, so all that remains is to determine the oxidation number of Cr, which we call y. The dichromate ion can be written as

$$[Cr_2^{(y)}O_7^{(2-)}]^{2-}$$

so that $$2(y) + 7(-2) = -2$$

or $$y = +6$$

Solution 4:

Solution From the molar mass of glucose, we write

$$4.07\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ mol\ C_6H_{12}O_6}{180.2\ \cancel{g\ C_6H_{12}O_6}} = 2.259 \times 10^{-2}\ mol\ C_6H_{12}O_6$$

Next, we calculate the volume of the solution that contains 2.259×10^{-2} mol of the solute. Rearranging Equation (4.2) gives

$$\begin{aligned} V &= \frac{n}{M} \\ &= \frac{2.259 \times 10^{-2}\ mol\ C_6H_{12}O_6}{3.16\ mol\ C_6H_{12}O_6/\ L\ soln} \times \frac{1000\ mL\ soln}{1\ L\ soln} \\ &= \boxed{7.15\ mL\ soln} \end{aligned}$$

Solution 5:

$$\text{Molarity} = \frac{0.715\ \text{mol glycine}}{495\ \cancel{\text{mL soln}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \text{L}} = \boxed{1.44\ M\ \text{glycine}}$$

Solution 6:

SOLUTION Writing the balanced equation:

$$Mg(OH)_2(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + 2H_2O(l)$$

Converting from grams of $Mg(OH)_2$ to moles:

$$\text{Moles of } Mg(OH)_2 = 0.10 \cancel{\text{g } Mg(OH)_2} \times \frac{1 \text{ mol } Mg(OH)_2}{58.33 \cancel{\text{g } Mg(OH)_2}} = 1.7\times10^{-3} \text{ mol } Mg(OH)_2$$

Converting from moles of $Mg(OH)_2$ to moles of HCl:

$$\text{Moles of HCl} = 1.7\times10^{-3} \cancel{\text{mol } Mg(OH)_2} \times \frac{2 \text{ mol HCl}}{1 \cancel{\text{mol } Mg(OH)_2}} = 3.4\times10^{-3} \text{ mol HCl}$$

Converting from moles of HCl to liters:

$$\text{Volume (L) of HCl} = 3.4\times10^{-3} \cancel{\text{mol HCl}} \times \frac{1 \text{ L}}{0.10 \cancel{\text{mol HCl}}} = \boxed{3.4\times10^{-2} \text{ L}}$$

Solution 7:

(a) Write the molecular equation

$$2KF(aq) + Sr(NO_3)_2(aq) \longrightarrow SrF_2(s) + 2KNO_3(aq)$$

Writing the total ionic equation:

$$2K^+(aq) + 2F^-(aq) + Sr^{2+}(aq) + 2NO_3^-(aq) \longrightarrow SrF_2(s) + 2K^+(aq) + 2NO_3^-(aq)$$

Writing the net ionic equation:

$$Sr^{2+}(aq) + 2F^-(aq) \longrightarrow SrF_2(s)$$

The spectator ions are K^+ and NO_3^-.

(b) The other ion combinations are ammonium bromide and sodium perchlorate. It shows that all ammonium, sodium, and most perchlorate compounds are soluble, and all bromides are soluble except those of Ag+, Pb^{2+}, Cu^+, and Hg_2^{2+}. Therefore, ***no*** reaction occurs. The compounds remain dissociated in solution as solvated ions.

Solution 8:

SOLUTION Writing the balanced equation:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Finding volume (L) of NaOH solution added:

$$\text{Volume (L) of solution} = (33.87\ \cancel{\text{mL}}\ \text{soln} - 0.55\ \cancel{\text{mL}}\ \text{soln}) \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}}$$

$$= 0.03332\ \text{L soln}$$

Finding amount (mol) of NaOH added:

$$\text{Moles of NaOH} = 0.03332\ \cancel{\text{L soln}} \times \frac{0.1524\ \text{mol NaOH}}{1\ \cancel{\text{L soln}}}$$

$$= 5.078 \times 10^{-3}\ \text{mol NaOH}$$

Finding amount (mol) of HCl originally present: Since the molar ratio is 1/1,

$$\text{Moles of HCl} = 5.078 \times 10^{-3}\ \cancel{\text{mol NaOH}} \times \frac{1\ \text{mol HCl}}{1\ \cancel{\text{mol NaOH}}} = 5.078 \times 10^{-3}\ \text{mol HCl}$$

Calculating molarity of HCl:

$$\text{Molarity of HCl} = \frac{5.078 \times 10^{-3}\ \text{mol HCl}}{50.00\ \cancel{\text{mL}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \text{L}}$$

$$= 0.1016\ M\ \text{HCl}$$

Solution 9:

SOLUTION **(a)** Assigning oxidation numbers:

$$2\overset{0}{Al}(s) + 3\overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4}(aq) \longrightarrow \overset{+3}{Al_2}(\overset{+6}{S}\overset{-2}{O_4})_3(aq) + 3\overset{0}{H_2}(g)$$

The O.N. of Al increased from 0 to +3 (Al lost electrons), so Al was oxidized;

Al is the reducing agent.

The O.N. of H decreased from +1 to 0 (H gained electrons), so H^+ was reduced;

H_2SO_4 is the oxidizing agent.

(b) Assigning oxidation numbers:

$$\overset{+2}{Pb}\overset{-2}{O}(s) + \overset{+2}{C}\overset{-2}{O}(g) \longrightarrow \overset{0}{Pb}(s) + \overset{+4}{C}\overset{-2}{O_2}(g)$$

Pb decreased its O.N. from +2 to 0, so PbO was reduced; PbO is the oxidizing agent.

C increased its O.N. from +2 to +4, so CO was oxidized; CO is the reducing agent.

In general, when a substance (such as CO) becomes one with more O atoms (such as CO_2), it is oxidized; and when a substance (such as PbO) becomes one with fewer O atoms (such as Pb), it is reduced.

(c) Assigning oxidation numbers:

$$\overset{0}{2H_2(g)} + \overset{0}{O_2(g)} \longrightarrow 2\overset{+1}{H_2}\overset{-2}{O}(g)$$

O_2 was reduced (O.N. of O decreased from 0 to -2); O_2 is the oxidizing agent.
H_2 was oxidized (O.N. of H increased from 0 to $+1$); H_2 is the reducing agent.
Oxygen is always the oxidizing agent in a combustion reaction.

Solution 10:

SOLUTION

1. Divide into half-reactions.

$MnO_4^- \longrightarrow MnO_2$ $\qquad$ $C_2O_4^{2-} \longrightarrow CO_3^{2-}$

2. Balance.

a. Atoms other than O and H,
Not needed

b. O atoms with H_2O,
$MnO_4^- \longrightarrow MnO_2 + 2H_2O$

c. H atoms with H^+,
$4H^+ + MnO_4^- \longrightarrow MnO_2 + 2H_2O$

d. Charge with e^-,
$3e^- + 4H^+ + MnO_4^- \longrightarrow MnO_2 + 2H_2O$
[reduction]

a. Atoms other than O and H,
$C_2O_4^{2-} \longrightarrow 2CO_3^{2-}$

b. O atoms with H_2O,
$2H_2O + C_2O_4^{2-} \longrightarrow 2CO_3^{2-}$

c. H atoms with H^+,
$2H_2O + C_2O_4^{2-} \longrightarrow 2CO_3^{2-} + 4H^+$

d. Charge with e^-,
$2H_2O + C_2O_4^{2-} \longrightarrow 2CO_3^{2-} + 4H^+ + 2e^-$
[oxidation]

3. Multiply each half-reaction, if necessary, by some integer to make e^- lost equal e^- gained.

$2(3e^- + 4H^+ + MnO_4^- \longrightarrow MnO_2 + 2H_2O)$
$6e^- + 8H^+ + 2MnO_4^- \longrightarrow 2MnO_2 + 4H_2O$

$3(2H_2O + C_2O_4^{2-} \longrightarrow 2CO_3^{2-} + 4H^+ + 2e^-)$
$6H_2O + 3C_2O_4^{2-} \longrightarrow 6CO_3^{2-} + 12H^+ + 6e^-$

4. Add half-reactions, and cancel substances appearing on both sides.
The six e^- cancel, eight H^+ cancel to leave four H^+ on the right, and four H_2O cancel to leave two H_2O on the left:

$$\cancel{6e^-} + \cancel{8H^+} + 2MnO_4^- \longrightarrow 2MnO_2 + \cancel{4H_2O}$$
$$2\,\cancel{6H_2O} + 3C_2O_4^{2-} \longrightarrow 6CO_3^{2-} + 4\,\cancel{12H^+} + \cancel{6e^-}$$
$$2MnO_4^- + 2H_2O + 3C_2O_4^{2-} \longrightarrow 2MnO_2 + 6CO_3^{2-} + 4H^+$$

4 Basic. Add OH^- to both sides to neutralize H^+, and cancel H_2O.
Adding four OH^- to both sides forms four H_2O on the right, two of which cancel the two H_2O on the left, leaving two H_2O on the right:

$$2MnO_4^- + 2H_2O + 3C_2O_4^{2-} + 4OH^- \longrightarrow 2MnO_2 + 6CO_3^{2-} + [4H^+ + 4OH^-]$$
$$2MnO_4^- + \cancel{2H_2O} + 3C_2O_4^{2-} + 4OH^- \longrightarrow 2MnO_2 + 6CO_3^{2-} + 2\,\cancel{4}H_2O$$

Including states of matter gives the final balanced equation:

$$2MnO_4^-(aq) + 3C_2O_4^{2-}(aq) + 4OH^-(aq) \longrightarrow 2MnO_2(s) + 6CO_3^{2-}(aq) + 2H_2O(l)$$

5. Check that atoms and charges balance.

(2Mn, 24O, 6C, 4H; 12−) $\longrightarrow$ (2Mn, 24O, 6C, 4H; 12−)

Solution 11:

Step 1. Divide the reaction into half-reactions, each of which contains the oxidized and reduced forms of one species. The two chromium species make up one half-reaction, and the two iodine species make up the other:

$$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$$
$$I^- \longrightarrow I_2$$

Step 2. Balance atoms and charges in each half-reaction. We use H_2O to balance O atoms, H^+ to balance H atoms, and e^- to balance positive charges.

- For the $Cr_2O_7^{2-}/Cr^{3+}$ half-reaction:
  a. *Balance atoms other than O and H.* We balance the two Cr on the left with a coefficient 2 on the right:
  $$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$$
  b. *Balance O atoms by adding H_2O molecules.* Each H_2O has one O atom, so we add seven H_2O on the right to balance the seven O in $Cr_2O_7^{2-}$:
  $$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$
  c. *Balance H atoms by adding H^+ ions.* Each H_2O contains two H, and we added seven H_2O, so we add 14 H^+ ions on the left:
  $$14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$
  d. *Balance charge by adding electrons.* Each H^+ ion has a 1+ charge, and 14 H^+ plus $Cr_2O_7^{2-}$ gives 12+ on the left. Two Cr^{3+} give 6+ on the right. There is an excess of 6+ on the left, so we add six e^- on the left:
  $$6e^- + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

  This half-reaction is balanced, and we see it is the *reduction* because electrons appear on the *left, as reactants:* the reactant $Cr_2O_7^{2-}$ gains electrons (is reduced), so $Cr_2O_7^{2-}$ is the *oxidizing agent.* (Note that the O.N. of Cr decreases from +6 on the left to +3 on the right.)

  For the I^-/I_2 half-reaction:
  a. *Balance atoms other than O and H.* Two I atoms on the right require a coefficient 2 on the left:
  $$2I^- \longrightarrow I_2$$
  b. *Balance O atoms with H_2O.* Not needed; there are no O atoms.
  c. *Balance H atoms with H^+.* Not needed; there are no H atoms.
  d. *Balance charge with e^-.* To balance the 2− on the left, we add two e^- on the right:
  $$2I^- \longrightarrow I_2 + 2e^-$$

  This half-reaction is balanced, and it is the *oxidation* because electrons appear on the *right, as products:* the reactant I^- loses electrons (is oxidized), so I^- is the *reducing agent.* (Note that the O.N. of I increases from −1 to 0.)

Step 3. Multiply each half-reaction, if necessary, by an integer so that the number of e^- lost in the oxidation equals the number of e^- gained in the reduction. Two e^- are lost in the oxidation and six e^- are gained in the reduction, so we multiply the oxidation by 3:

$$3(2I^- \longrightarrow I_2 + 2e^-)$$
$$6I^- \longrightarrow 3I_2 + 6e^-$$

Step 4. Add the half-reactions together, canceling substances that appear on both sides, and include states of matter. In this example, only the electrons cancel:

$$\cancel{6e^-} + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$
$$6I^- \longrightarrow 3I_2 + \cancel{6e^-}$$
$$\overline{6I^-(aq) + 14H^+(aq) + Cr_2O_7^{2-}(aq) \longrightarrow 3I_2(s) + 7H_2O(l) + 2Cr^{3+}(aq)}$$

Step 5. Check that atoms and charges balance:

Reactants (6I, 14H, 2Cr, 7O; 6+) $\longrightarrow$ products (6I, 14H, 2Cr, 7O; 6+)