



## THE COPPERBELT UNIVERSITY

### DEPARTMENT OF CHEMISTRY

#### CH110: TUTORIAL SHEET 2

#### REACTIONS IN AQUEOUS SOLUTIONS

##### Term I (2023-2024)

1. The compounds  $\text{K}_2\text{CO}_3$ ,  $\text{Na}_2\text{CO}_3$ ,  $\text{KCl}$  and  $\text{NaCl}$  are soluble in water, but  $\text{CaCO}_3$  is not. Given the following molecular equation:  
$$\text{CaCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{NaCl}(\text{aq})$$
  - (a) Write *molecular equation* for the reaction between solutions of  $\text{CaCl}_2$  and  $\text{K}_2\text{CO}_3$
  - (b) Write the *ionic equation* for the reaction between solutions of  $\text{CaCl}_2$  and  $\text{K}_2\text{CO}_3$
  - (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  $\text{CaCl}_2$  and  $\text{K}_2\text{CO}_3$
2. What volume of 16 M sulfuric acid must be used to prepare 1.5 L of a 0.10-M  $\text{H}_2\text{SO}_4$  solution?
3. Assign oxidation numbers to all the elements in the following compounds and ion:  
(a)  $\text{Na}_2\text{O}$ , (b)  $\text{HNO}_2$ , (c)  $\text{Cr}_2\text{O}_7^{2-}$
4. Define molarity. Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous  $\text{HCl}$  in enough water to make 26.8 mL of solution.
5. Glycine ( $\text{H}_2\text{NCH}_2\text{COOH}$ ) 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  $\text{HCl}$  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,  $\text{Mg}(\text{OH})_2$ , which reacts with the acid to form water and magnesium chloride solution. As a government chemist testing commercial antacids, you use 0.10 M  $\text{HCl}$  to simulate the acid concentration in the

stomach. How many liters of “stomach acid” react with a tablet containing 0.10 g of  $\text{Mg}(\text{OH})_2$ ?

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)  $\rightarrow$

(b) Ammonium perchlorate(aq) + sodium bromide(aq)  $\rightarrow$

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?

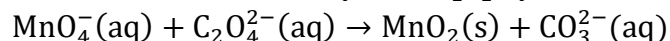
9. Identify the oxidizing agent and reducing agent in each of the following:

(a)  $2\text{Al}(\text{s}) + 3\text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Al}_2(\text{SO}_4)_3(\text{aq}) + 3\text{H}_2(\text{g})$

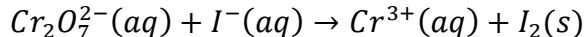
(b)  $\text{PbO}(\text{s}) + \text{CO}(\text{g}) \rightarrow \text{Pb}(\text{s}) + \text{CO}_2(\text{g})$

(c)  $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

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  $\text{NaMnO}_4$  and  $\text{Na}_2\text{C}_2\text{O}_4$  in basic solution:



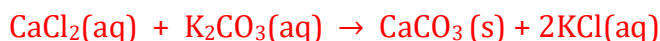
11. Balance the redox reaction between dichromate ion and iodide ion to form chromium(III) ion and solid iodine, which occurs in acidic solution.



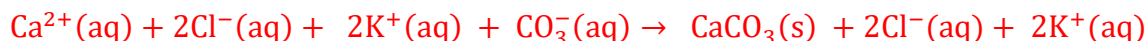
## Solutions

### Solution 1:

- (a) Write molecular equation for the reaction between solutions of  $\text{CaCl}_2$  and  $\text{K}_2\text{CO}_3$



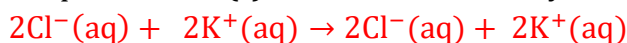
- (b) Write the ionic equation for the reaction between solutions of  $\text{CaCl}_2$  and  $\text{K}_2\text{CO}_3$



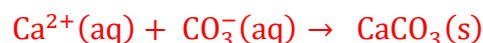
- (c) What is meant by the term spectator ion

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



- (e) Write the net ionic equation for the reaction between solutions of  $\text{CaCl}_2$  and  $\text{K}_2\text{CO}_3$



### Solution 2:

#### Solution

*Where are we going?*

To find the volume of  $\text{H}_2\text{SO}_4$  required to prepare the solution

*What do we know?*

- › 1.5 L of 0.10 M  $\text{H}_2\text{SO}_4$  is required
- › We have 16 M  $\text{H}_2\text{SO}_4$

*What information do we need to find the volume of  $\text{H}_2\text{SO}_4$ ?*

- › Moles of  $\text{H}_2\text{SO}_4$  in the required solution

*How do we get there?*

*What are the moles of  $\text{H}_2\text{SO}_4$  required?*

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

$$1.5 \text{ L solution} \times \frac{0.10 \text{ mol H}_2\text{SO}_4}{\text{L solution}} = 0.15 \text{ mol H}_2\text{SO}_4$$

*What volume of 16 M  $\text{H}_2\text{SO}_4$  contains 0.15 mole of  $\text{H}_2\text{SO}_4$ ?*

$$V \times \frac{16 \text{ mol H}_2\text{SO}_4}{\text{L solution}} = 0.15 \text{ mol H}_2\text{SO}_4$$

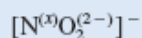
Solving for  $V$  gives

$$V = \frac{0.15 \text{ mol H}_2\text{SO}_4}{\frac{16 \text{ mol H}_2\text{SO}_4}{1 \text{ L solution}}} = 9.4 \times 10^{-3} \text{ L or } 9.4 \text{ mL solution}$$

### Solution 3:

#### Solution

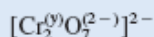
- (a) By rule 2, we see that sodium has an oxidation number of +1 ( $\text{Na}^+$ ) and oxygen's oxidation number is  $-2$  ( $\text{O}^{2-}$ ).
- (b) This is the formula for nitrous acid, which yields a  $\text{H}^+$  ion and a  $\text{NO}_2^-$  ion in solution. From rule 4, we see that H has an oxidation 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



$$\text{so that} \quad x + 2(-2) = -1$$

$$\text{or} \quad x = +3$$

- (c) From rule 6, we see that the sum of the oxidation numbers in the dichromate ion  $\text{Cr}_2\text{O}_7^{2-}$  must be  $-2$ . We know that the oxidation 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



$$\text{so that} \quad 2(y) + 7(-2) = -2$$

$$\text{or} \quad y = +6$$

### Solution 4:

**Solution** From the molar mass of glucose, we write

$$4.07 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6} \times \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.2 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6}} = 2.259 \times 10^{-2} \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6$$

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

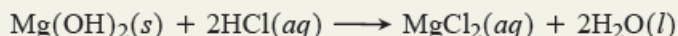
$$\begin{aligned} V &= \frac{n}{M} \\ &= \frac{2.259 \times 10^{-2} \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{3.16 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6 / \text{L soln}} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} \\ &= 7.15 \text{ mL soln} \end{aligned}$$

### Solution 5:

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

### Solution 6:

**SOLUTION** Writing the balanced equation:



Converting from grams of  $\text{Mg}(\text{OH})_2$  to moles:

$$\text{Moles of } \text{Mg}(\text{OH})_2 = 0.10 \text{ g } \text{Mg}(\text{OH})_2 \times \frac{1 \text{ mol } \text{Mg}(\text{OH})_2}{58.33 \text{ g } \text{Mg}(\text{OH})_2} = 1.7 \times 10^{-3} \text{ mol } \text{Mg}(\text{OH})_2$$

Converting from moles of  $\text{Mg}(\text{OH})_2$  to moles of  $\text{HCl}$ :

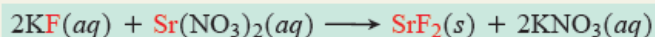
$$\text{Moles of } \text{HCl} = 1.7 \times 10^{-3} \text{ mol } \text{Mg}(\text{OH})_2 \times \frac{2 \text{ mol } \text{HCl}}{1 \text{ mol } \text{Mg}(\text{OH})_2} = 3.4 \times 10^{-3} \text{ mol } \text{HCl}$$

Converting from moles of  $\text{HCl}$  to liters:

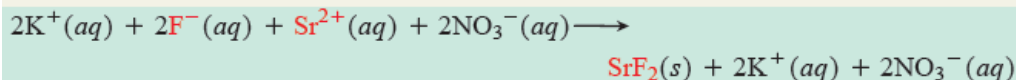
$$\text{Volume (L) of } \text{HCl} = 3.4 \times 10^{-3} \text{ mol } \text{HCl} \times \frac{1 \text{ L}}{0.10 \text{ mol } \text{HCl}} = 3.4 \times 10^{-2} \text{ L}$$

### Solution 7:

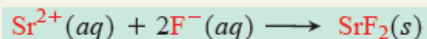
(a) Write the molecular equation



Writing the total ionic equation:



Writing the net ionic equation:



The spectator ions are  $\text{K}^+$  and  $\text{NO}_3^-$ .

(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  $\text{Ag}^+$ ,  $\text{Pb}^{2+}$ ,  $\text{Cu}^+$ , and  $\text{Hg}_2^{2+}$ . Therefore, **no** reaction occurs. The compounds remain dissociated in solution as solvated ions.

**Solution 8:****SOLUTION** Writing the balanced equation:

Finding volume (L) of NaOH solution added:

$$\begin{aligned}\text{Volume (L) of solution} &= (33.87 \text{ mL soln} - 0.55 \text{ mL soln}) \times \frac{1 \text{ L}}{1000 \text{ mL}} \\ &= 0.03332 \text{ L soln}\end{aligned}$$

Finding amount (mol) of NaOH added:

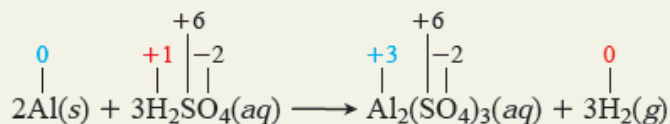
$$\begin{aligned}\text{Moles of NaOH} &= 0.03332 \text{ L soln} \times \frac{0.1524 \text{ mol NaOH}}{1 \text{ L soln}} \\ &= 5.078 \times 10^{-3} \text{ mol NaOH}\end{aligned}$$

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

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

Calculating molarity of HCl:

$$\begin{aligned}\text{Molarity of HCl} &= \frac{5.078 \times 10^{-3} \text{ mol HCl}}{50.00 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \\ &= 0.1016 \text{ M HCl}\end{aligned}$$

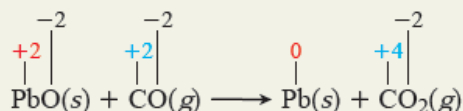
**Solution 9:****SOLUTION** (a) Assigning oxidation numbers:

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  $\text{H}^+$  was reduced; $\text{H}_2\text{SO}_4$  is the oxidizing agent.

(b) Assigning oxidation numbers:

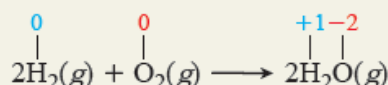


Pb decreased its O.N. from +2 to 0, so PbO was reduced; PbO is the oxidizing agent.

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  $\text{CO}_2$ ), 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:



O<sub>2</sub> was reduced (O.N. of O decreased from 0 to -2); O<sub>2</sub> is the oxidizing agent.

H<sub>2</sub> was oxidized (O.N. of H increased from 0 to +1); H<sub>2</sub> is the reducing agent.

Oxygen is always the oxidizing agent in a combustion reaction.

## Solution 10:

### SOLUTION

1. Divide into half-reactions.

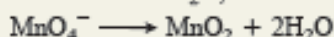


2. Balance.

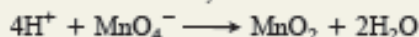
a. Atoms other than O and H,

Not needed

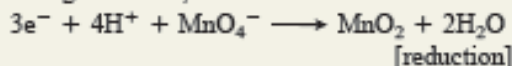
b. O atoms with H<sub>2</sub>O,



c. H atoms with H<sup>+</sup>,



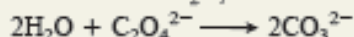
d. Charge with e<sup>-</sup>,



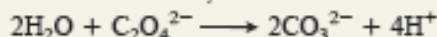
a. Atoms other than O and H,



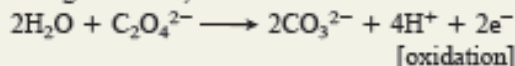
b. O atoms with H<sub>2</sub>O,



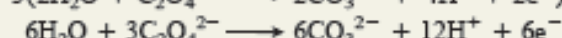
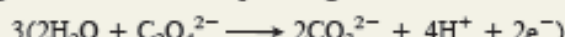
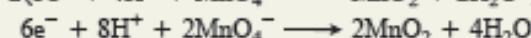
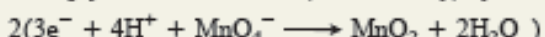
c. H atoms with H<sup>+</sup>,



d. Charge with e<sup>-</sup>,

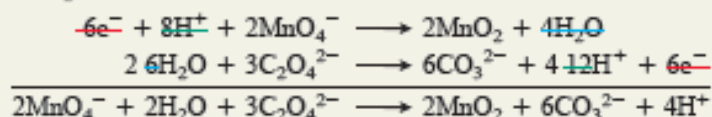


3. Multiply each half-reaction, if necessary, by some integer to make e<sup>-</sup> lost equal e<sup>-</sup> gained.



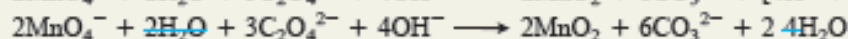
4. Add half-reactions, and cancel substances appearing on both sides.

The six e<sup>-</sup> cancel, eight H<sup>+</sup> cancel to leave four H<sup>+</sup> on the right, and four H<sub>2</sub>O cancel to leave two H<sub>2</sub>O on the left:

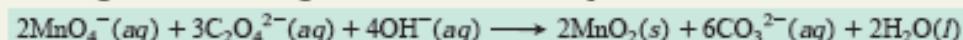


4 Basic. Add OH<sup>-</sup> to both sides to neutralize H<sup>+</sup>, and cancel H<sub>2</sub>O.

Adding four OH<sup>-</sup> to both sides forms four H<sub>2</sub>O on the right, two of which cancel the two H<sub>2</sub>O on the left, leaving two H<sub>2</sub>O on the right:



Including states of matter gives the final balanced equation:

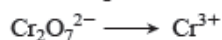


5. Check that atoms and charges balance.



### 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:*



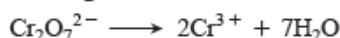
*Step 2. Balance atoms and charges in each half-reaction. We use  $\text{H}_2\text{O}$  to balance O atoms,  $\text{H}^+$  to balance H atoms, and  $\text{e}^-$  to balance positive charges.*

- For the  $\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}$  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:



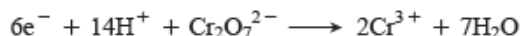
- b. *Balance O atoms by adding  $\text{H}_2\text{O}$  molecules.* Each  $\text{H}_2\text{O}$  has one O atom, so we add seven  $\text{H}_2\text{O}$  on the right to balance the seven O in  $\text{Cr}_2\text{O}_7^{2-}$ :



- c. *Balance H atoms by adding  $\text{H}^+$  ions.* Each  $\text{H}_2\text{O}$  contains two H, and we added seven  $\text{H}_2\text{O}$ , so we add 14  $\text{H}^+$  ions on the left:



- d. *Balance charge by adding electrons.* Each  $\text{H}^+$  ion has a 1+ charge, and 14  $\text{H}^+$  plus  $\text{Cr}_2\text{O}_7^{2-}$  gives 12+ on the left. Two  $\text{Cr}^{3+}$  give 6+ on the right. There is an excess of 6+ on the left, so we add six  $\text{e}^-$  on the left:



This half-reaction is balanced, and we see it is the *reduction* because electrons appear on the *left, as reactants*: the reactant  $\text{Cr}_2\text{O}_7^{2-}$  gains electrons (is reduced), so  $\text{Cr}_2\text{O}_7^{2-}$  is the *oxidizing agent*. (Note that the O.N. of Cr decreases from +6 on the left to +3 on the right.)

For the  $\text{I}^-/\text{I}_2$  half-reaction:

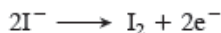
- a. *Balance atoms other than O and H.* Two I atoms on the right require a coefficient 2 on the left:



- b. *Balance O atoms with  $\text{H}_2\text{O}$ .* Not needed; there are no O atoms.

- c. *Balance H atoms with  $\text{H}^+$ .* Not needed; there are no H atoms.

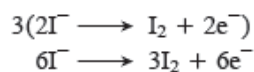
- d. *Balance charge with  $\text{e}^-$ .* To balance the 2- on the left, we add two  $\text{e}^-$  on the right:



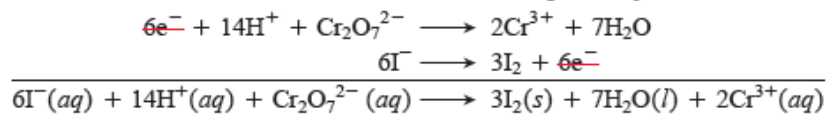
This half-reaction is balanced, and it is the *oxidation* because electrons appear on the *right, as products*: the reactant  $\text{I}^-$  loses electrons (is oxidized), so  $\text{I}^-$  is the *reducing agent*. (Note that the O.N. of I increases from -1 to 0.)



*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:*



*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:*



*Step 5. Check that atoms and charges balance:*

