Friday Worksheet **Name: ……………….**

Revision Unit 3 and 4 worksheet 1

1) Iron is an essential element and has many functions throughout the body. Many factors can deplete the body of iron. Factors such as bleeding or poor diet. A certain brand of iron supplement contains soluble iron(II) salts that replenish Fe^{2+} in the body in case of deficiency. Iron supplements usually contain about 25 mg per tablet of iron(II). The permanganate ion (MnO_4^-) is converted to Mn^{2+} in the presence of Fe^{2+}, while the Fe(II) is converted to Fe(III).

Potassium permanganate ($KMnO_4$), has a molar mass of 158.0 atomic mass units and is a good primary standard, To confirm the amount, in grams, of Fe^{2+} ions in each tablet a student followed the experimental procedure outlined below.

i. Weigh a tablet in a beaker and record its mass
ii. Add 100 ml of deionised water and heat until the tablet is dissolved.
iii. Allow the solution to cool and transfer it to a 250 ml volumetric flask and make up to the mark with deionised water.
iv. Prepare a standard $KMnO_4$ solution by accurately weighing out approximately 1.58 grams of potassium permanganate and adding it to approximately 100 mL of 2.00 M H_2SO_4. This is then transferred to a 1000 ml volumetric flask and made up to the mark with deionised water.
v. Using a 20 mL pipette transfer 20.0 mL from the 1000 mL volumetric flask into another 1000 mL volumetric flask and make up to the mark with deionised water . Label it "Solution2"
vi. Pipette a 25.0 ml aliquot of the standard $KMnO_4$ solution from the volumetric flask labelled "Solution 2" into a conical flask.
vii. Fill a burette with the unknown iron(II) solution from the 250 mL volumetric flask. Record the starting volume of the burette.
viii. Add the unknown Fe(II) solution from the burette into the conical flask containing the 25.0 mL standard $KMnO_4$ until the purple colour of the $KMnO_4$ solution disappears. Record the final reading from the burette.

Results from a typical experiment are shown below:

- Mass of potassium permanganate used: 1.63 g
- Titration results:

titre	initial/ml ± 0.05	final/ml ± 0.05	volume/ml ± 0.1
1	0.00	17.86	17.86

a) What makes potassium permanganate a good primary standard?
It is available to a high degree of purity, it has a large relative mass and is stable in the environment.

b) Calculate the concentration of the standard $KMnO_4$ solution in the flask labelled "Solution 2"

potassium permanganate solution in the original 1000 mL volumetric flask=

1.63/158.034 = 1.03×10^{-2}M

potassium permanganate solution in the 1000 mL volumetric flask labelled "Solution 2"
Since this is a dilution we can use the formula
$C_1V_1 = C_2V_2$
Where
$C_1 = 1.03 \times 10^{-2}$ (The concentration in the original 1000 mL volumetric flask)
V_1 = 0.0200 L (The volume delivered from the original volumetric flask to the one labelled "Solution 2)
V_2 = 1.00 L (The final volume of the solution)
C_2 = ?
=> $C_2 = C_1V_1 / V_2 = 1.03 \times 10^{-2} \times 0.0200 / 1.00 = 2.06 \times 10^{-4}$

c) Circle the type of reaction that takes place between the MnO_4^- and the Fe(II) and justify your answer.

- Acid/base

- Redox oxidation => $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e$
reduction => $MnO_4^-(aq) + 8H^+(aq) + 5e \rightarrow Mn^{2+}(aq) + 2H_2O(l)$

- Esterification.

d) Write the overall reaction that occurs.

$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 2H_2O(l) + 5Fe^{3+}(aq)$

Calculate the mol of MnO_4^- present in the conical flask
mol of MnO_4^- = V X C = $0.0250 \times 2.06 \times 10^{-4} = 5.16 \times 10^{-6}$

e) Calculate the amount of Fe(II), in grams, in the 17.86 mL titre.
$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 2H_2O(l) + 5Fe^{3+}(aq)$
According to the stoichiometric ratio for every mol of MnO_4^- that reacts 5 mol of Fe(II) also react.
mol of Fe(II) = $5.16 \times 10^{-6} \times 5 = 2.58 \times 10^{-5}$
Calculate the mass in mg of Fe(II) present in the 17.86 titre.
mass = mol X atomic mass of iron = $2.58 \times 10^{-5} \times 55.9 = 1.44 \times 10^{-3}$ g

f) Calculate the mass of iron in one capsule.
Since 1.44×10^{-3} grams of iron, as found in in e) above, came from 17.86 mL To find the amount, in grams, of iron that came from one capsule
=> (250/17.86) X 1.44×10^{-3} = 0.0201 g = 20.1 mg

g) What are some improvements that the student can make to their investigation procedure?

- Test more than one capsule.
- Perform the titration of each capsule more than once until concordant results are obtained.

h) Another student conducted the same investigation at the titration results are shown below. Complete the table and give the average titre.

titre	initial/ml ± 0.05	final/ml ± 0.05	volume/ml ± 0.1
1	1.00	19.20	18.20
2	19.20	37.38	18.18
3	1.35	19.55	18.20
4	19.55	38.20	18.65
5	4.42	23.62	19.20

Use only concordant result to calculate the average.
Results that differ by no more than 0.10 mL from highest to lowest value are considered concordant.
Hence (18.20 + 18.18 + 18.10) / 3 = 18.16 mL

2) Sodium is produced from the electrolysis of sodium salts. Which of the following would be the best choice for the electrolyte and the anode in a commercial cell? Explain

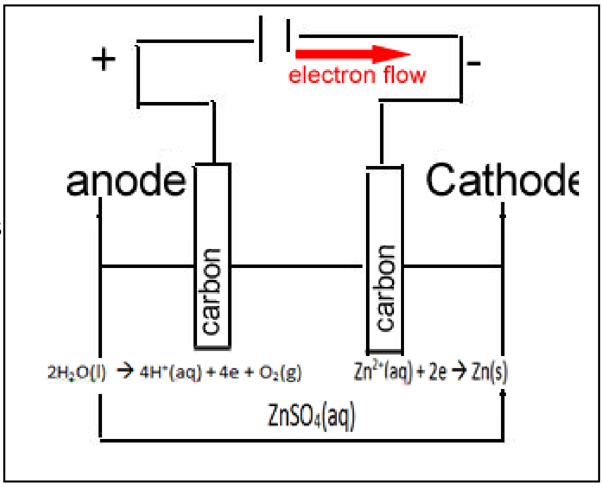
a) 0.1M NaCl solution, using iron electrodes
b) Molten NaCl using zinc electrodes
c) 1.0 M NaCl solution using carbon electrodes
d) Molten NaCl using carbon electrodes

Sodium cannot be produced from a solution. H_2O is a stronger oxidant than Na^+ and hence will react at the cathode in preference to Na^+.
Option b) cannot use the zinc electrode. Zinc metal is a stronger reductant than Cl^- ions and will be oxidised at the anode according to the equation
$Zn(s) \rightarrow Zn^{2+}(l) + 2e$

3) A classroom experiment was set up to simulate the industrial extraction of zinc metal from an aqueous solution of zinc ions by electrolysis. In this experiment 200 mL of 1.50 M $ZnSO_4$ solution was electrolysed at 25°C using inert carbon electrodes.

a) Draw a diagram of the electrolytic cell. Label the following.

i. direction of electron flow
ii. cathode
iii. anode
iv. polarity of each electrode
v. oxidation half-equation
vi. reduction half-equation
vii. material that each electrode is made from



b) A mass of 1.900 g of zinc is produced in 65.0 minutes. Calculate the electric current, in Amps, supplied to the cell during the electrolysis. Express your answer to an appropriate number of significant figures.

Step 1 find the mol of zinc deposited
=> 1.900 /65.4 = 0.0291
Step 2 find the charge needed to deposit this amount of Zn.
=> Zn^{2+} (aq) + 2e → Zn(s)
Find the mol of electrons
=> 0.0291 X 2 = 0.0582
Step 3 find the charge this amount of electrons represents
=> 0.0582 X 96500 = 5616 C
Step 3 find the current
=> 5616 C = It
=> 5616 / (65.0 X 60) = 1.44 Amps

4) What is the pH of a 252 mL sample of a 0.230 M Propanoic acid, solution?

From the data book obtain the value of the Ka of propanoic acid.
Step 1 write the Ka expression
=> Ka = $[H_3O^+][C_2H_5COOH] / [C_2H_5COOH] = 1.3 \times 10^{-5}$.
Step 2 Assume negligible amount of the acid is ionised
=> Hence $[C_2H_5COOH] = 0.230$
Step 3 Since $[H_3O^+] = [C_2H_5COOH]$
=> $[H_3O^+]^2 / 0.230 = 1.3 \times 10^{-5}$
Step 4 find the $[H_3O^+]$
=> $[H_3O^+]^2 = 0.230 \times 1.3 \times 10^{-5}$
=> $[H_3O^+] = 10^{-2.762}$
Step 5 find the pH
2.76