RPS DEGREE COLLEGE BALANA (MAHENDERGARH)-123029



Lab Manual

Chemistry (B.Sc.3rd& 4thSemester)

Department of Chemistry

PHYSICAL CHEMISTRY

Semester 3rd

Chemical kinetics

- 1. To determine the specific reaction rates of hydrolysis of ethyl/methyl ester catalyzed by hydrogen ion at room temperature.
- 2. Refractive Index

2To determine the Refractive Index if given liquid and calculation of specific and molar refractivity.

3.Electrochemistry

3.Determination of p^{H} of a solution.

Semester 4th

Chemical kinetics

1. To study the effect of acid strength on hydrolysis of ester.

Electrochemistry

2. Strength of acid by p^H measurement.

Refractive Index

3. Determination of concentration of binary mixture by measurement of refractive Index.

Electrochemical Cell

4. Setting of a Galvanic Cell and determination of cell voltage.

Semester -3

Experiment :1

Aim:

To determine the specific reaction rates of hydrolysis of ethyl/methyl ester catalyzed by hydrogen ion at room temperature.

PRINCIPLE:

The hydrolysis of an ester occurs according to the equation

 $CH_{3}COOC_{2}H_{5} + H_{2}^{O} \longrightarrow CH_{3}COOH + C_{2}H_{5}OH$

This reaction follows pseudo first order kinetics.

PROCEDURE:

100 ml of 0.5 N HCl is taken in a clean dry conical flask. 5 ml of ester is pipetted out into the conical flask and the mixture is immediately withdrawn into another dry conical flask. A stop watch is started simultaneously. The reaction is then arrested by the addition of ice cubes and the mixture is titrated against 0.2 N NaOH using phenolphthalein as indicator. End point is the appearance of permanent pink colour. The volume of NaOH consumed in this titration is taken as V_0 .

5 ml of acid – ester mixture is similarly withdrawn after 10, 20, 30, ..., 60 minutes respectively and titrated against NaOH using phenolphthalein as indicator. The volume of NaOH consumed for each of the above time intervals (t), is taken as V_t.

The contents are transferred into boiling tube with a cap and heated in a water bath for about 15 minutes. 5 ml of this mixture is withdrawn and titrated against NaOH to get V_{∞} .

CALCULATION:

The rate constant K is determined using the equation,

$$K = \frac{2.303}{t} \log \frac{(V_{\infty} - V_0)}{(V_{\infty} - V_t)}$$

Rate constant is also determined graphically by plotting

 $\log \left(\boldsymbol{V}_{\infty} - \boldsymbol{V}_{t} \right) \qquad \text{Vs time.}$

OBSERVATIO:

S.No.	Time	Volume				
	Min	of	$(V_{\infty} - V_t)$	$log(V_{\infty}-V_t)$	$2.303 (V_{\infty} - V_{0})$	
		NaOH	ml	ml	$K = \frac{1}{t} \log \frac{1}{(V_{\infty} - V_t)}$	
		ml			(K1) min-1	
1	0					
2	10					
3	20					
4	30					
5	40					
6	50					
7	60					
8	8					

RESULT:

The Rate Constant for the hydrolysis of an ester from

- 1. Calculated value =
- 2. Graphical value =

Rate=k[ester]

EXPERIMENT – 2

AIM:-To determine theRefractive Index if given liquid and calculation of specific and molar refractivity.

APPARATUS REQUIRED:-

Pyknometer or specific gravity bottle, abde's refractometer, source of light, dropper, weight box.

CHEMICAL REQUIRED:-

Given liquid whose specific refractivity is to be determined.

PROCEDURE:-

1. Determination of density of given liquid

- i. Clean the pyknometer with chromic acid (solid $K_2Cr_2O_7$ + conc. H_2SO_4) carefully and then wash it thoroughly several times with water. Rinse it with ethyl alcohol and dry it with the help of a drier.
- ii. Suspend the pyknometer from the end of the balance beam by means of a copper hook and weigh it accurately.
- iii. Now attach a clean rubber tube to the end 'a' of pyknometer and immerse the end 'b' in distilled water and suck the distilled water through the rubber tube gently till water fills the bulb and stands to the mark on end a. In case water stands below the mark then suck more water and in case the water stands above the mark then remove excess water with the help of a filter paper strip. Ensure that no air bubble is present in pyknometer. Dry it from outside by wiping with a filter paper and weigh it.
- iv. Then remove the water and dry it again. Repeat the experiment with the given liquid.
- 2. To determine the refractivity of the given liquid.

- i. Open the prism box, clean the prism surfaces with ethyl alcohol (with soft cotton) and allow it to dry.
- ii. Introduce 3-4 drops of the given liquid with the help of a dropper between the prisms and press them tightly together.
- iii. Allow the light from the lamp to fall on the mirror. Set the mirror to reflect maximum light to the prism. Black spots in the field of view are due to insufficient liquid, then introduce 1 or 2drops of given liquid more.
- iv. Rotate the prism box by moving the lever until the boundary between the shaded and bright parts appears sharp.
- v. If the light shade disc has a band of colours, make it sharp by rotating the knob of compensator.
- vi. Adjust the prism box lever with screw so that light shade disc passes through the centre of the cross wires.
- vii. Read the refractive index of the given liquid directly on the scale through the eye-piece. Take three readings from bright to dark field and another three readings from dark to bright field. Take average of these readings.

OBSERVATIONS:-

Weight of empty pyknometer $= W_1 g$

Weight of pyknometer + water = $w_2 g$

Weight of pyknometer + given liquid = $w_3 g$

$$\therefore Weight of water = (w_2 -w_1) g Weight of liquid = (w_3 - w_1) g$$

: Density of liquid = $\frac{weig \ ht \ of \ liquid}{weig \ ht \ of \ water} = \frac{w_3 - w_1}{w_2 - w_1} g$

Assuming density of water as one we can calculate the specific refractivity of given liquid

:. Specific refractivity,
$$\mathbf{r} = \left(\frac{n_r^2 - 1}{n_r^2 + 2}\right) \frac{1}{d}$$

Where n_r is refractive index and d is density of given liquid.

RESULT:-

Therefore specific refractivity of given liquid is 'r'.

Table: Refractive indeces of Some liquids temperatures

Table 1. Density of water at different

Liquid	Refractive
	index
Acetic acid	1.3718
Acetone	1.3616
Benzene	1.5044
CCl_4	1.4631
Chloroform	1.4486
Ethyl lcohol	1.3620
Ethyl acetate	1.3726
Methyl	1.3312
alcohol	1.4999
Toluene	1.3333
Water	

Temperature in	Density in g/ml
°C	
0	0.99987
1	0.99993
2	0.99997
3	0.99999
4	1.00000
5	0.99999
б	0.99997
7	0.99993
8	0.99988
9	0.99981
10	0.99973
11	0.99963
12	0.99952
13	0.99940
14	0.99927
15	0.99913
16	0.99897
17	0.99884
18	0.99862
19	0.99843
20	0.99823

21	0.99802
22	0.99780
23	0.99757
24	0.99733
25	0.99708
26	0.99681
27	0.99654
28	0.99626
29	0.99598
30	0.99568
31	0.99537
32	0.99506
33	0.99473
34	0.99440
35	0.99406
36	0.99372
37	0399336
38	0.99300
39	0.99262
40	0.99225
50	0.98807

Experiment:3

AIM:-.Determination of p^H of a solution.

THEORY:-

Poteniometric method is applicable to wide range of titration, provided in indicator electrode is available an indicator electrode is the whose potential indicance the change in the concentration of ion to be titrated. The variation of potential of an electrode with the concentration of ion with which it is in equilibrium is used as an indicator in volumemetric analysis. It is not possible to

determine the electrode potential separately, hence the indicator electrode is used in conjunction with a reference electrode, in which potential at one of the reference electrode remains constant during the course of titration e.g. saturated calomel electrode (SCE) while potential of other electrode depends on the activity of H^+ ion in the solution.

A solution of acid is taken in a beaker

Ag, AgCl/1MHCl || KCl aq.|| calomel electrode

CHEMICAL USED:-

- 1. Approx 0.1 N HCl solution
- 2. quinhydrone
- 3. KCl salt bridge
- 4. Pure H₂

APPARATUS USED:-

- 1. Potentiometer
- 2. Glass electrode
- 3. Calomel electrode
- 4. Beaker
- 5. Stirrer
- 6. Burette

PROCEDURE:-

- I. Set all the apparatus properly.
- II. Take 20 ml of acid solution in a beaker and platinum electrode is dipped into the solution to saturate the solution mix a quinhydrone.
- III. Connect two electrodes to a potentiometer.
- IV. Measure the e.m.f. of the cell.
- V. Repeat the experiment 3 times.
- VI. Note down the e.m.f. after each time.

OBSERVATION:-

Sr.No.	EMF		

	(Volts)		
1.	E1		
2.	E ₂		
3.	E ₃		
4.			
5.			
6.	_		
7.	_		
8.	_		

Mean emf=-----

CALCULATION:-

 E_{cell} =-0.059 pH

Result-pH of the solution is-

Semester -4

EXPERIMENT:-1

AIM:

To determine the rate constant of the hydrolysis of Ethyl acetate using an acid as a catalyst.

PRINCIPLE:

The hydrolysis of an ester occurs according to the equation

$$CH_{3}COOC_{2}H_{5} + H_{2}^{O} \longrightarrow CH_{3}COOH + C_{2}H_{5}OH$$

This reaction follows pseudo first order kinetics.

PROCEDURE:

100 ml of 0.5 N HCl is taken in a clean dry conical flask. 5 ml of ester is pipetted out into the conical flask and the mixture is immediately withdrawn into another dry conical flask. A stop watch is started simultaneously. The reaction is then arrested by the addition of ice cubes and the mixture is titrated against 0.2 N NaOH using phenolphthalein as indicator. End point is the appearance of permanent pink colour. The volume of NaOH consumed in this titration is taken as V_0 .

5 ml of acid – ester mixture is similarly withdrawn after 10, 20, 30, ..., 60 minutes respectively and titrated against NaOH using phenolphthalein as indicator. The volume of NaOH consumed for each of the above time intervals (t), is taken as V_t.

The contents are transferred into boiling tube with a cap and heated in a water bath for about 15 minutes. 5 ml of this mixture is withdrawn and titrated against NaOH to get V_{∞} .

CALCULATION:

The rate constant K is determined using the equation,

$$K = \frac{2.303}{t} \log \frac{(V_{\infty} - V_0)}{(V_{\infty} - V_t)}$$

Rate constant is also determined graphically by plotting $\log (V_{\infty} - V_t)$

 (V_t) Vs time.

OBSERVATIO:

S.No.	Time Min	Volume of NaOH ml	$(V_{\infty} - V_t)$ ml	$log(V_{\infty} - V_t)$ ml	$K = \frac{2.303}{t} \log \frac{(V_{\infty} - V_{0})}{(V_{\infty} - V_{t})}$ (K1) min-1	$K = \frac{2.303}{t} \log \frac{(V_{\infty} - V_0)}{(V_{\infty} - V_t)}$ (K2) min-1
1	0					
2	10					

3	20			
4	30			
5	40			
6	50			
7	60			
8	8			

RESULT:

The Rate Constant for the hydrolysis of an ester from

- 1. Calculated value =
- 2. Graphical value =

$$\frac{K1}{K2} =$$

Experiment 2

AIM:-Strength of acid by p^H measurement

THEORY:-

In pH metric titrations, when some alkali is added to an acid there will be an decrease in concentration of H^+ ion and increase in pH of the solution.

pH is the negative logarithm of hydrogen ion or hydronium ion concentration.

 $pH = -log[H_3O^+)$

In the beginning, the pH increases slowly but at the end point, the increase is very large due to rapid increase in the fraction of the H^+ ion removed by the addition of constant volume of the alkali. After the end point, pH increases slowly. A plot of pH vs volume of alkali added will give the equivalence point.

CHEMICAL USED:-

- 1. Standard alkali or $\frac{N}{10}$ NaOH
- 2. HCl solution
- 3. A buffer solution of known pH (.05 M potassium hydrogen phthalate, pH = 4)
- 4. Distilled water

APPARATUS USED:-

- 1. pH meter
- 2. Glass and calomel electrodes
- 3. Burette
- 4. Beaker

PROCEDURE:-

- 1. Wash the electrodes with distilled water.
- 2. Take 20 ml of given acid solution in a beaker and dip the electrodes in it.
- 3. Measure the pH of the solution using previously calibrated pH meter with a glass electrode.
- 4. Mix 0.5 ml of NaOH solution from a burette each time and measure the pH after stirring the solution.
- 5. Mix NaOH solution dropwise near the end point, and measure the pH after each mixing.
- **6.** Repeat the experiment to take some more readings beyond the equivalent point.

OBSERVATION:-

Volume of given HCl taken in beaker = 20 ml

Normality of NaOH solution = $\frac{N}{20}$

Volume of NaOH used for complete neutralization or to get end point = V ml

CALCULATION:-

By ploting the pH vs volume of alkali added, find the equivalence point Using normality equation,

$$N_1V_1 = N_2V_2$$
(HCI) (NaOH)
$$N_1 \times 20 = \frac{N}{20} \times V$$

 \therefore N₁ is Normality of HClsolution $=\frac{V}{400}$

Strength of HClsolution = Normality × Eq. wt.

$$=\frac{N}{400}$$
 × 36.5 = x g/litre

RESULT:-

Strength of given HClsolution = x g/litre PRECAUTIONS:-

- 1. When electrodes not in use, should be placed in distilled water.
- 2. With buffer solution of known pH.pH meter should be calibrated.
- 3. After each a mixing solution should be stirred.
- 4. Near the end point, NaOH solution should be mixed dropwise.

EXPERIMENT – 3

AIM:-To determine the specific refractivity of a given binary mixtutre.

APPARATUS REQUIRED:-

Pyknometer or specific gravity bottle, abde's refractometer, source of light, dropper, weight box.

CHEMICAL REQUIRED:-

Given mixture of two liquid whose specific refractivity is to be determined.

PROCEDURE:-

3. Determination of density of given liquid

- v. Clean the pyknometer with chromic acid (solid $K_2Cr_2O_7$ + conc. H_2SO_{4}) carefullyand then wash it thoroughly several times with water. Rinse it with ethyl alcohol and dry it with the help of a drier.
- vi. Suspend the pyknometer from the end of the balance beam by means of a copper hook and weigh it accurately.
- vii. Now attach a clean rubber tube to the end 'a' of pyknometer and immerse the end 'b' in distilled water and suck the distilled water through the rubber tube gently till water fills the bulb and stands to the mark on end a. In case water stands below the mark then suck more water and in case the water stands above the mark then remove excess water with the help of a filter paper strip. Ensure that no air bubble is present in pyknometer. Dry it from outside by wiping with a filter paper and weigh it.
- viii. Then remove the water and dry it again. Repeat the experiment with the given liquid.

4. To determine the refractivity of the given liquid mixture.

- viii. Open the prism box, clean the prism surfaces with ethyl alcohol (with soft cotton) and allow it to dry.
- ix. Introduce 3-4 drops of the given liquid mixture with the help of a dropper between the prisms and press them tightly together.
- x. Allow the light from the lamp to fall on the mirror. Set the mirror to reflect maximum light to the prism. Black spots in the field of view are due to insufficient liquid, then introduce 1 or 2drops of given liquid more.
- xi. Rotate the prism box by moving the lever until the boundary between the shaded and bright parts appears sharp.
- xii. If the light shade disc has a band of colours, make it sharp by rotating the knob of compensator.
- xiii. Adjust the prism box lever with screw so that light shade disc passes through the centre of the cross wires.
- xiv. Read the refractive index of the given liquid mixture directly on the scale through the eye-piece. Take three readings from bright to dark field and another three readings from dark to bright field. Take average of these readings.

OBSERVATIONS:-

Weight of empty pyknometer $= W_1 g$

Weight of pyknometer + water = $w_2 g$

Weight of pyknometer + given liquid mixture= $w_3 g$

: Density of liquid mixture = $\frac{weig ht of liquid}{weig ht of water} = \frac{w_3 - w_1}{w_2 - w_1} g$

Assuming density of water as one we can calculate the specific refractivity of given liquid

$$\therefore \qquad \text{Specific refractivity, } \mathbf{r} = \left(\frac{n_r^2 - 1}{n_r^2 + 2}\right) \frac{1}{d}$$

Where n_r is refractive index and d is density of given liquid.

RESULT:-

Therefore specific refractivity of given liquid is 'r'.

Table: Refractive indeces of Some liquids temperatures Table 1. Density of water at different

Liquid	Refractive index
Acetic acid	1.3718
Acetone	1.3616

Benzene	1.5044
CCl_4	1.4631
Chloroform	1.4486
Ethyl lcohol	1.3620
Ethyl acetate	1.3726
Methyl	1.3312
alcohol	1.4999
Toluene	1.3333
Water	

Temperature in	Density in g/ml
°C	
0	0.99987
1	0.99993
2	0.99997
3	0.99999
4	1.00000
5	0.99999
6	0.99997
7	0.99993
8	0.99988
9	0.99981
10	0.99973
11	0.99963
12	0.99952
13	0.99940
14	0.99927
15	0.99913
16	0.99897
17	0.99884
18	0.99862
19	0.99843
20	0.99823
21	0.99802
22	0.99780
23	0.99757
24	0.99733
25	0.99708
26	0.99681
27	0.99654
28	0.99626
29	0.99598
30	0.99568
31	0.99537

32	0.99506
33	0.99473
34	0.99440
35	0.99406
36	0.99372
37	0399336
38	0.99300
39	0.99262
40	0.99225
50	0.98807

Experiment -4

Aim:

Setting of a Galvanic Cell and determination of cell voltage.

Introduction:

Chemical reactions involving the transfer of electrons from one reactant to another are called oxidation-reduction reactions or redox reactions. In a redox reaction, two half-reactions occur; one reactant gives up electrons (undergoes oxidation) and another reactant gains electrons (undergoes reduction). A piece of zinc going into a solution as zinc ions, with each Zn atom giving up 2 electrons, is an example of an oxidation half-reaction.

 $Zn(s) \rightarrow Zn2+(aq) + 2e$ - (1)

The oxidation number of Zn(s) is 0 and the oxidation number of the Zn2+ is +2. Therefore, in this half-reaction, the oxidation number increases, which is another way of defining an oxidation. In contrast, the reverse reaction, in which Zn2+ ions gain 2 electrons to become Zn atoms, is an example of reduction.

 $Zn2+(aq) + 2e \rightarrow Zn(s)$ (2)

In a reduction there is a decrease (or reduction) in oxidation number. Chemical equation representing half-reactions must be both mass and charge balanced. In the half-reactions above, there is one zinc on both sides of the equation. The charge is balanced because the 2+ charge on the zinc ion is balanced by two electrons, 2e-, giving zero net charge on both sides.

Another example of reduction is the formation of solid copper from copper ions in solution.

 $Cu2+(aq) + 2e \rightarrow Cu(s)$ (3)

In this half-reaction the oxidation number of the aqueous copper is +2, which decreases to 0 for the solid copper, and again charge and mass are balanced. However, no half-reaction can occur by itself. A redox reaction results when an oxidation and a reduction half-reaction are combined to complete a transfer of electrons as in the following example:

$$Zn(s) + Cu2 + (aq) \rightarrow Zn2 + (aq) + Cu(s)$$
 (4)

The electrons are not shown because they are neither reactants nor products but have simply been transferred from one species to another (from Zn to Cu2+ in this case). In this redox reaction, the Zn(s) is referred to as the reducing agent because it causes the Cu2+ to be reduced to Cu. The Cu2+ is called the oxidizing agent because it causes the Zn(s) to be oxidized to Zn2+.

Any half-reaction can be expressed as a reduction as illustrated in the case where equation (1) can be reversed to equation (2). A measure of the tendency for a reduction to occur is its reduction potential, E, measured in units of volts. At standard conditions, 25 °C and concentrations of 1.0 M for the aqueous ions, the measured voltage of the reduction halfreaction is defined as the standard reduction potential, E° .

A galvanic cell or voltaic cell is a device in which a redox reaction, such as the one in equation (4), spontaneously occurs and produces an electric current. In order for the transfer of electrons in a redox reaction to produce an electric current and be useful, the electrons are made to pass through an external electrically conducting wire instead of being directly transferred between the oxidizing and reducing agents. The design of a galvanic cell (shown in Figure 1 for the equation (4) reaction) allows this to occur. In a galvanic cell, two solutions, one containing the ions of the oxidation half-reaction and the other containing the ions of the reduction half-reaction, are placed in separated compartments called half-cells. For each half-cell, the metal, which is called an electrode, is placed in the solution and connected to an external wire. The electrode at which oxidation occurs is called the anode [Zn in equation (4)] and the electrode at which reduction occurs is called the cathode [Cu in equation (4)]. The two half-cells are connected by a salt-bridge that allows a "current" of ions from one half-cell to the other to complete the circuit of electron current in the external wires. When the two electrodes are connected to an electric load (such as a light bulb or voltmeter) the circuit is completed, the oxidation-reduction reaction occurs, and electrons move from the anode (-) to the cathode (+), producing an electric current.

The cell potential, Ecell, which is a measure of the voltage that the battery can provide, is calculated from the half-cell reduction potentials:

Ecell = Ecathode - Eanode

At standard conditions, indicated by the superscript o, the standard cell potential, E° cell, is based upon the standard reduction potentials, as shown in equation (5).

 E° cell = E° cathode – E° anode (5)

Based on the values for the standard reduction potentials for the two half-cells in equation (4) [-0.76 V for zinc anode and +0.34 V for copper cathode], the standard cell potential, E°cell, for the galvanic cell in Figure 1 would be:

 $E^{\circ}cell = +0.34 V - (-0.76 V) = +1.10 V$

The positive voltage for Eocell indicates that at standard conditions the reaction is spontaneous. Recall that $\Delta Go = -nFEocell$, so that a positive Eocell results in a negative ΔGo . Thus the redox reaction in equation (4) would produce an electric current when set up as a galvanic cell.

art A. Redox Reactions: In this experiment you will observe several redox reactions in which metals are placed in solutions containing different metal ions. From your observations you will determine whether a redox reaction is occurring and write balanced redox equations for any that occur. For example, since as shown above, Cu2+ has a greater tendency to be reduced than Zn2+, you would

expect that placing Zn metal into a solution of Cu2+ ions would result in a direct redox reaction. Cu2+ ions are reduced to Cu metal which is deposited on the Zn metal surface, while the Zn metal is oxidized to Zn 2+ ions which go into the solution. The redox equation for this reaction is therefore equation (4).

Part B: Reduction Potentials: You will then construct a series of three galvanic cells combining the zinc half-reaction with three different metal half-reactions (Cu, Fe and Pb). You will measure the cell potentials, E°cell, using a Vernier voltage probe as shown in Figure 3. You will use 1.0 M solutions for both half-cells, so Q = 1 and $\ln Q = 0$ for the reaction. Thus the cell potential measured will be the same as E°cell as evident from the Nernst equation (6). You will then use your

measuredE°cell values, the known zinc standard reduction potential, $E^{\circ} = -0.76 \text{ V}$, and equation (5) to calculate the E° values for the three different half-reactions.

Part C: Nernst Equation for varying Cu2+ concentrations: Galvanic cells with different known Cu2+ concentrations and a fixed Zn2+ concentration will be prepared and their cell potentials measured. A plot of the Nernst equation, Ecell vs. ln[Cu2+], constructed from this data will then be used to find [Cu2+] of an unknown solution.

Part D: Determine the E° for a voltaic cell using Cu and unknown metal: Finally, you will measure the potential of a voltaic cell combining an unknown metal electrode with Cu (E° = 0.34 V). By measurement of the cell potential and use of equation (5), you will identify the unknown metal from its calculated value of E°. The unknown will have a more negative E° than Cu, so the Cu will have a greater tendency to be reduced and thus will be the cathode when the E°cell is positive.

Part A. Redox Reactions: 1. Fill four cells in each of four columns of a 24-well microcell plate about three-fourths full with 1.0 M Cu(NO3)2, 1.0 M FeSO4, 1.0 M Pb(NO3)2, and 1.0 M Zn(NO3)2 as shown in the diagram below. 2. Polish small strips (4 each) of Cu, Fe, Pb and Zn with steel wool or sand paper and place them on a paper towel with written labels to insure that the metals are not mixed up with each other. Partially submerge the strips into the cell rows as shown below. Place only part of the metal into the solution so that any sign of a reaction (such as deposit of a metal on the submerged part of the strip) can be determined by comparison with the unsubmerged portion of the metal strip.

Cu(NO3)2 FeSO4 Pb(NO3)2 Zn(NO3)2 Cu(s) Fe(s) Pb(s) Zn(s)

3. After 5 minutes examine each cell carefully to see if any metal displacement redox reaction has occurred. Record your observations in a table as shown above in your laboratory notebook. Then repolish and rinse the strips and return them to labeled paper towels located at the rear of the lab. If you are unsure about the identity of any strips, give them to your lab instructor.

Part B: Reduction Potentials: 1. Before constructing the galvanic cells, set up the Vernier system in DATAMATE with the voltage probe connected to channel 1 of the LabPro interface. 2. Hit CLEAR on the main screen of DATAMATE and the program will check for sensors. Once the voltage probe has been identified, and a reading near 0 volts is shown, you are ready to record data from the main screen. Obtain a 1.5 V battery from the instructor's desk and connect the red lead of the voltage probe to the (+) end of the battery (the cathode) and the black lead to the (-) end (the anode). If the voltage reading is not 1.5 V +/- 0.2 V, inform you instructor. 3. Wet a 2–3 inch long strip of filter paper with KNO3 solution. Remove excess liquid gently (the paper easily rips when wet!) by blotting it on a paper towel and fold the paper into a U-shape; this will serve as your salt bridge. Place the salt bridge so that it will be immersed into each of two solutions in adjacent wells of a 12-well microcell plate (see Figure 3 below). 4. Construct a galvanic cell by adding solutions of 1.0 M Cu(NO3)2 as the aqueous Cu2+ and 1.0 M Zn(NO3)2 as the aqueous Zn2+ to the two adjacent wells. Do not immerse the

copper and zinc at this time. Connect the leads from the voltage probe to the zinc and copper electrodes. Recall that the red lead must be connected to the cathode and the black to the anode to obtain a positive cell voltage. 5. Immerse the electrodes in the solutions and remove them in 5-10 seconds while avoiding contact with the salt bridge. If the voltage reading on the Vernier calculator is positive, the electrodes are connected correctly; if not, repolish and rinse the electrodes and reverse the connections.

To Vernier

Filter paper salt bridge

Microcell plate

6. Record the positive cell potential, in volts, in your laboratory notebook. You will find that the voltage recorded is less than what you would expect based on standard reduction potentials. For example the Zn/Cu cell may be less than the 1.10 V calculated in the Introduction section. This is partly due to the presence of an oxide formed on one of the electrodes, a process which occurs very rapidly for easily oxidized metals and which changes the half-cell potential. 7. Repeat steps 3-6 for galvanic cells of Fe(in 1.0 M FeSO4) - Zn and Pb(in 1.0 M Pb(NO3)2) - Zn. Use a freshly prepared filter paper strip for the salt bridge of each cell.

Part C: Nernst Equation for varying Cu2+ concentrations: 1. Prepare three dilute solutions of CuSO4 by serial dilution from a 1.0 M CuSO4 stock solution as follows: Transfer1.0 mL of 1.0 M CuSO4 stock solution into a labeled (with tape) 100.00 mL volumetric flask and dilute to the mark with de-ionized water to form a 0.01 M solution. Rinse the pipet with a small volume of the solution to be transferred prior to use. Next take the 0.01 M solution just formed and transfer 1.0 mL of it into a labeled 100.00 mL volumetric flask. Dilute to the mark with de-ionized water to form the 0.0001 M solution. Repeat one more time so that you have the following set of solutions in labeled flasks: Solution A - 1 M CuSO4

(stock solution) Solution B – 0.01 M CuSO4 Solution C – 0.0001 M CuSO4 Solution D – 0.000001 M CuSO4 2. Prepare a half-cell of Cu2+ by placing the copper solution D into a 12-microcell well. 3. Remove the copper and zinc electrodes from the half-cells used previously and clean and re-polish them. 4. Place 1.0 M Zn(NO3)2 in a cell next to the Cu2+ half-cell. Connect the two halfcells with a freshly prepared salt bridge. Connect the copper and zinc electrodes to the correct voltage probe leads. Measure and record the cell potential in your laboratory notebook using the same technique (5-10 second immersion) with the voltage probe as in Part B. 5. Remove the CuSO4 solution with a disposable pipet and repeat the measurement of Ecell (steps 3-4) for the remaining three copper solutions in order of increasing concentration. Then, in the same way, measure the Ecell for the unknown Cu2+ solution. Record the measured values in a table as shown below in your notebook.

Solution D C B A Unknown [Cu2+] 0.000001 M 0.0001 M 0.01 M 1.0 M ? Volts

Part D: Determine the E° for a voltaic cell using Cu and unknown metal:1 1. Obtain a small amount of the unknown electrolyte solution labeled "1.0 M X ion" and the corresponding metal strip, "X". This metal is one of the metals in the table of Standard Reduction Potentials at the end of the In-Lab section. 2. Use a disposable pipet to transfer a small amount of 1.0 M X ion solution to a well adjacent to the 1.0 M CuSO4 solution in a 12-microcell-test plate. 3. Make a new salt bridge by soaking a short length of filter paper in the KNO3 solution. 4. Connect the X and Cu half-cells with the filter paper. Measure the positive potential of the X-Cu voltaic cell using the same technique as in Part B (Red lead to the Cu, which is the cathode).

5. After recording the potential once (5-10 seconds), remove both electrodes from the solutions and clean and polish each electrode. Set up the galvanic cell again. Connect the voltage probe as before. 6. Record the potential again. If the two measured potentials do not agree within .1 volts, test the galvanic cell a third time and record the potential immediately after making the connection with the voltage probe. Calculate the average of the measured potentials to use in the Post-Lab calculation.

Standard Reduction Potentials:

Electrode Eo

 $\begin{array}{l} Ag++e-\rightarrow Ag +0.80 \ V \ Cu2++2e-\rightarrow Cu +0.34 \ V \ Pb2++2e-\rightarrow Pb \ -0.13V \\ Fe2++2e-\rightarrow Fe \ -0.44 \ V \ Zn2++2e-\rightarrow Zn \ -0.76 \ V \ Al3++3e-\rightarrow Al \ -1.66 \ V \\ Mg2++2e-\rightarrow Mg \ -2.37 \ V \end{array}$

Part A. Redox Reactions: For each of the metals, indicate the redox reactions you observed in a table as shown below. Write NR for (no reaction) where none was observed. 10 Cu(NO3)2 FeSO4 Pb(NO3)2 Zn(NO3)2 Cu(s) :

Fe(s):

Pb(s):

Zn(s): ** Explain your observations by calculating the Eo for each reaction observed using equation (5) and values for the standard reduction potentials of the metals to determine why certain reactions are spontaneous (** e.g., Cu2+ + Zn \rightarrow Cu + Zn2+, E0cell = 1.10V) 10

Part B: Reduction Potentials: Report the cell potential for each galvanic cell and state which electrode corresponds to the cathode and which to the anode. Given Eo = -0.76 V for the Zn2+ / Zn half-cell, and your measured Eocell, calculate the reduction potential at the Cu, Fe, and Pb cathodes 15 Galvanic cell Ecell , Volts Cathode halfreaction Anode halfreaction Reduction Potential, volts Cu(s) - Zn(s): Fe(s) - Zn(s): Pb(s) - Zn(s): 1. Show a sample calculation for the case of the Fe(s) - Zn(s) cell. 5

Part C: Nernst Equation: 1. Construct and attach to your report a graph (similar to that shown in Figure 2) of the measured values of Ecell versus $\ln[Cu2+]$ from your data table. 5 2. Find the best fit line for the plot and state: "The equation of the line is _____". 5 3. From the equation for the line, determine [Cu2+] of the unknown solution. State: "The [Cu2+] of the unknown solution is _____."

Part D: Determine the E° for a voltaic cell using Cu and an unknown metal:1 1.Use equation (5) to solve for E° anode of the unknown metal, keeping in mind that the Cu is the cathode for the cell when the E° cell is positive. Show the calculation of E° anode . 10 2. Compare the E° anode of the unknown to the values listed in the table given at the end of the In-Lab section and state: "The unknown metal is ." 10