APPENDIX A. SUPPLEMENTAL MATERIAL

A.1 THE QUADRATIC FORMULA

The solutions to many problems involving chemical equilibria (and other phenomena) result in a quadratic equation of the form:

$$ax^2 + bx + c = 0 \qquad \text{(A-1)}$$

Thus extensive use of the quadratic formula will be made in this text to obtain a final answer to such problems

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \qquad \text{(A-2)}$$

It should first be noted that x will always have two values (the equation has two roots). Both values are legitimate mathematically, but only *one* has real meaning chemically in the problems which will be encountered. The challenge is to select the correct root. Often this is obvious, *e.g.* if x represents a concentration and one of the roots is negative, the negative root is chemically meaningless, since all concentrations must be positive.

It should also be noted that, although imaginary roots are mathematically acceptable, they are not acceptable in the type of problems encountered in this text. Thus if $(b^2\text{-}4ac) < 0$ (negative), something is WRONG, go back over the problem and correct the mistake.

Often the quadratic formula can be simplified. Three separate cases will be examined.

1) If $b^2 << -4ac$ (either a or c, but not both, must be negative), then

$$\sqrt{b^2 - 4ac} \cong 2\sqrt{-ac} \qquad \text{(A-3)}$$

Thus

$$x \cong \frac{-b}{2a} \pm \sqrt{\frac{-c}{a}} \qquad \text{(A-4)}$$

2) If $b^2 \approx -4ac$, then no simplification can be made and the full formula must be used.

3) If $b^2 >> -4ac$, then

$$\sqrt{b^2 - 4ac} \cong b \qquad \text{(A-5)}$$

and

$$x \cong \frac{-b \pm b}{2a} \qquad \text{(A-6)}$$

which gives the two roots:

$$x \approx -b/a \quad \text{and} \quad x \approx 0 \qquad \text{(A-7)}$$

If the desired root is $-b/a$, there is no problem. However, if the other root is desired, the answer that $x \approx 0$ is not satisfactory. All this indicates is that x is small, but not how small. One might then decide to use the full equation, and this often gives the desired result. Sometimes, however, b and $(b^2\text{-}4ac)^{1/2}$ are so close to each other that the calculator or computer being used does not carry enough significant figures to perform the calculation correctly. All is not lost, since it can be shown that the other root is approximately $-c/b$ (given without proof). Thus

$$x \approx -b/a \quad \text{and} \quad x \approx -c/b \qquad \text{(A-7a)}$$

A.2 THE CORRECT USE OF CALCULATORS

Many students round their answers after every step in a multi-step calculation. This is incorrect and can lead to rounding errors, besides it is easier and quicker to proceed from one intermediate step to the next without changing any of the intermediate results. For example assume that you wish to perform the following calculation:

$y = (1.376 + 4.28)/8.5 - 0.801$

If this is done without rounding the result is -0.13558. This would be rounded to -0.14 since there should be two significant figures in the result according to the usual rules of thumb. However, if rounding is done at each step, the following results:

$(1.376 + 4.28) = 5.656$ or 5.66 after rounding

$5.66/8.5 = 0.66588$ or 0.67 after rounding

$0.67 - 0.801 = -0.131$ or -0.13 after rounding.

It is obvious that the two methods produce different results. The preferred method is to round only after the last calculation has been performed.

A.3 BALANCING REDOX REACTIONS

The hydrogen ion, in acidic solution, or the hydroxide ion, in basic solution, is often involved in redox reactions, and must be taken into account in the balancing of redox equations. Thus the pH of the solution is often specified when one is asked to balance a redox equation. The method recommended in this text is the half reaction method, since it fits in well with the other uses of half reactions in Chapter 5 on electrochemistry.

In order to successfully balance any redox reaction by any method, both redox couples must be known. These are then arranged in half reactions and each half reaction balanced separately. Finally the two half reactions are added to obtain the fully balanced reaction.

A.3.1 ACIDIC SOLUTIONS

Most individuals find that balancing redox reactions in acidic solution is easier than in basic solution. Thus acidic media will be examined first. You must know the correct reactants and products (both redox couples). As an example say that you know that dichromate ion will oxidize V(III) to vanadate ion and itself undergo reduction to Cr(III).

$$Cr_2O_7^{2-} + V^{3+} \rightarrow VO_3^- + Cr^{3+}$$

To balance this or any other redox reaction in acidic solution follow the steps outlined below:

STEP 1. Separate the reactants and products into the two redox couples and write them as half reactions (which will not be balanced at this time).

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

$$V^{3+} \rightarrow VO_3^-$$

STEP 2. Balance all elements except H and O in these two half reactions.

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

$$V^{3+} \rightarrow VO_3^-$$

STEP 3. For each half reaction determine which side of the reaction has excess oxygen. For each excess oxygen add two H^+ to the same side and add one H_2O to the opposite side. This will balance the H and O. In fact all elements in each half reaction should be balanced at this point.

$$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

$$V^{3+} + 3H_2O \rightarrow VO_3^- + 6H^+$$

STEP 4. Sum the charges on each side of each half reaction. Add electrons to the side of each half reaction which has the most positive charge so that the charges also balance. At this point you should have two balanced half reactions with electrons on opposite sides of the reactions. If this is not the case, you have made a mistake - check over your previous work.

$$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

12+ 6+ (add 6 electrons to the left)

$$V^{3+} + 3H_2O \rightarrow VO_3^- + 6H^+$$

3+ 5+ (add 2 electrons to the right)

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

$$V^{3+} + 3H_2O \rightarrow VO_3^- + 6H^+ + 2e^-$$

STEP 5. Multiply both half reactions by appropriate constants so that the number of electrons appearing in each equation is the same. Then add the two half reactions, cancelling the electrons, to obtain the final result. This is easily done for the above equations by multiplying the equation involving vanadium by three.

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

$$3V^{3+} + 9H_2O \rightarrow 3VO_3^- + 18H^+ + 6e^-$$

$$Cr_2O_7^{2-} + 14H^+ + 3V^{3+} + 9H_2O \rightarrow 2Cr^{3+} + 7H_2O + 3VO_3^- + 18H^+$$

or

$$Cr_2O_7^{2-} + 3V^{3+} + 2H_2O \rightarrow 2Cr^{3+} + 3VO_3^- + 4H^+$$

You should double check your answer by determining whether all elements and the charges balance in your final result. If not, go back and correct your mistake.

A.3.2 BASIC SOLUTIONS

The easiest way to balance redox equations in basic media is to proceed as if you were in acidic media. Follow the five steps described above. This will result in a balanced equation involving H^+ on one side of the equation. In STEP 6 add a number of OH^- to *both* sides of the full, balanced equation equal to the number of H^+ in the equation. You must add OH^- to both sides since this equation is already balanced, and you must keep it balanced. The OH^- on the side with the H^+ "reacts" to form H_2O, which usually cancels some water on the opposite side. After this cancellation, you should have a balanced equation involving OH^- and H_2O, in other words an equation for a basic solution. Again you should double check this for elemental and charge balance.

As an example take the reaction of permanganate ion with hydrogen peroxide to form manganese dioxide and oxygen.

$$MnO_4^- + H_2O_2 \rightarrow MnO_2(s) + O_2$$

To balance this equation proceed as follows:

STEPS 1 AND 2.

$$MnO_4^- \rightarrow MnO_2(s)$$

$$H_2O_2 \rightarrow O_2$$

STEP 3.

$$MnO_4^- + 4H^+ \rightarrow MnO_2(s) + 2H_2O$$

$$H_2O_2 \rightarrow O_2 + 2H^+$$

STEP 4.

$$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2(s) + 2H_2O$$

$$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$$

STEP 5.

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

$$3H_2O_2 \rightarrow 3O_2 + 6H^+ + 6e^-$$

$$2MnO_4^- + 8H^+ + 3H_2O_2 \rightarrow 2MnO_2(s) + 4H_2O + 3O_2 + 6H^+$$

or

$$2MnO_4^- + 2H^+ + 3H_2O_2 \rightarrow 2MnO_2(s) + 4H_2O + 3O_2$$

STEP 6. Since there are two H^+ on the left side of the above equation, add two OH^- to both sides.

$$2MnO_4^- + 2H^+ + 2OH^- + 3H_2O_2 \rightarrow 2MnO_2(s) + 4H_2O + 3O_2 + 2OH^-$$

"React" the H^+ and the OH^- to form water on the left and cancel this with the water on the right to get the final result.

$$2MnO_4^- + 2H_2O + 3H_2O_2 \rightarrow 2MnO_2(s) + 4H_2O + 3O_2 + 2OH^-$$

$$2MnO_4^- + 3H_2O_2 \rightarrow 2MnO_2(s) + 2H_2O + 3O_2 + 2OH^-$$

A.4 ELECTROCHEMICAL CALCULATIONS

A.4.1 DETERMINATION OF THE CELL VOLTAGE

The half cell method of calculating cell voltages is recommended for most analytical calculations. However, the full cell method may also be used, if desired. The first step in this latter method is to determine the cell standard potential as outlined in Chapter 5 on redox reactions. This value and the cell reagent concentrations are then substituted into the full Nernst equation and the cell voltage is calculated. This is fairly straight forward. However, several statements should be made in this regard. First, all reagents involved in the redox reactions must be included in the Nernst equation. Some individuals place only the two redox couples in this equation, neglecting ions like H^+. If H^+ is involved in the reaction, the cell voltage depends on its concentration and it *must* be included in the calculation. In fact H^+ is involved in many redox reactions and these are usually buffered to maintain its concentration constant. Second, calculation of the standard potential as suggested earlier guarantees that its value will be positive and that the reaction is written spontaneously under standard conditions. However, the calculated cell voltage under different conditions may be negative. If this occurs, it simply means that the reaction is written in the direction in which it will not proceed spontaneously. Simply reverse the reaction (and the sign of the calculated voltage). Third, on occasion the same substance will be present in both half cells, but at different concentrations. H^+ is a substance sometimes found under this circumstance. In this case, that ion will *not cancel or add* when the two half reactions are added to get the full cell reaction and must be left in two places in the full Nernst equation to obtain the correct cell voltage.

The example used to demonstrate this method will be the same as that used to demonstrate the half cell method. The cell consists of two platinum electrodes, one in a compartment containing I^- at 0.100 M and I_2 at 0.0100 M, the other in a compartment containing Cr(III) at 0.0100 M, $Cr_2O_7^{2-}$ at 0.0100 M and H^+ at 0.100 M. As was shown in Chapter 5, the full cell reaction and standard potential are as follows:

$$Cr_2O_7^{2-} + 6I^- + 14H^+ \rightarrow 2Cr^{3+} + 3I_2 + 7H_2O \qquad E° = 0.71\ V$$

Therefore the Nernst equation for this cell will be

$$E_{cell} = 0.71V - \frac{0.0592V}{6}\log\frac{[I_2]^3[Cr^{3+}]^2}{[Cr_2O_7^{2-}][I^-]^6[H^+]^{14}} \qquad \text{(A-8)}$$

On substituting the concentrations into Eqn. (A-8),

E_{cell} = 0.71 V - (0.00987 V)log$(0.01)^3(0.01)^2/(0.01)(0.1)^6(0.1)^{14}$

E_{cell} = 0.71 V - (0.00987 V)log$(1.00\text{x}10^{12})$ = 0.71 V - (0.00987)(12.00)

E_{cell} = 0.71 V - 0.118 V = 0.59 V

This value is the same as that calculated using the half cell method. Since the calculated cell voltage is positive, the reaction as written above is spontaneous. The electrode dipping into the iodine containing solution is the anode (negative), that in the chromium containing solution is the cathode (positive), also in agreement with the half cell method.

A.4.2 DETERMINATION OF K_e

Equilibrium calculations are usually performed in electrochemistry using the Nernst equation, not the law of mass action. However, on occasion it may be useful to know the equilibrium constant for a redox reaction. This is obtained from the standard potential as follows from the basic relations:

$$\Delta G° = -nFE° \qquad \text{(A-9)}$$

$$\Delta G° = -RT\ln K_e \qquad \text{(A-10)}$$

Combining Eqns. (A-9) and (A-10) gives

$$nFE° = RT\ln K_e \qquad \text{(A-11)}$$

or

$$K_e = \exp\{nFE°/RT\} \qquad \text{(A-12)}$$

Obviously, if the equilibrium constant is the known quantity, then the standard potential can be calculated. This in fact is one very useful method for obtaining standard potentials.

The equilibrium constant for the reaction used above in the example of cell voltage calculation using the full Nernst equation would be

K_e = exp{6(96490 coul/mol)(0.71 V)/(8.3143 V·coul/mol·K)(298.15 K)

K_e = exp(165.8) = $1.0\text{x}10^{72}$

APPENDIX B. TABLES

B.1 ATOMIC MASSES OF THE COMMON ELEMENTS

Element	Symbol	Atomic Mass	Element	Symbol	Atomic Mass
Aluminum	Al	26.982	Mercury	Hg	200.59
Antimony	Sb	121.75	Molybdenum	Mo	95.94
Argon	Ar	39.948	Neon	Ne	20.179
Arsenic	As	74.922	Nickel	Ni	58.71
Barium	Ba	137.33	Nitrogen	N	14.007
Beryllium	Be	9.012	Osmium	Os	190.2
Bismuth	Bi	208.980	Oxygen	O	15.999
Boron	B	10.811	Palladium	Pd	106.42
Bromine	Br	79.904	Phosphorus	P	30.974
Cadmium	Cd	112.41	Platinum	Pt	195.09
Calcium	Ca	40.08	Potassium	K	39.098
Carbon	C	12.011	Rhenium	Re	186.207
Cerium	Ce	140.12	Rhodium	Rh	102.906
Cesium	Cs	132.905	Rubidium	Rb	85.468
Chlorine	Cl	35.453	Ruthenium	Ru	101.07
Chromium	Cr	51.996	Selenium	Se	78.96
Cobalt	Co	58.933	Silicon	Si	28.086
Copper	Cu	63.546	Silver	Ag	107.868
Fluorine	F	18.998	Sodium	Na	22.990
Gallium	Ga	69.735	Strontium	Sr	87.62
Germanium	Ge	72.59	Sulfur	S	32.066
Gold	Au	196.967	Tantalum	Ta	180.948
Helium	He	4.003	Tellurium	Te	127.60
Hydrogen	H	1.0079	Thallium	Tl	204.37
Indium	In	114.82	Thorium	Th	232.038
Iodine	I	126.905	Tin	Sn	118.69
Iridium	Ir	192.22	Titanium	Ti	47.90
Iron	Fe	55.847	Tungsten	W	183.85
Krypton	Kr	83.80	Uranium	U	238.029
Lanthanum	La	138.906	Vanadium	V	50.942
Lead	Pb	207.2	Xenon	Xe	131.30
Lithium	Li	6.941	Zinc	Zn	65.38
Magnesium	Mg	24.305	Zirconium	Zr	91.224
Manganese	Mn	54.938			

B.2 VALUES OF t FOR VARIOUS LEVELS OF PROBABILITY

Degrees of Freedom	Confidence Interval 80%	90%	95%	98%	99%
1	3.078	6.314	12.706	31.821	63.657
2	1.886	2.920	4.303	6.965	9.925
3	1.638	2.353	3.182	4.541	5.841
4	1.533	2.132	2.776	3.747	4.604
5	1.476	2.015	2.571	3.365	4.032
6	1.440	1.943	2.447	3.143	3.707
7	1.415	1.895	2.365	2.998	3.499
8	1.397	1.860	2.306	2.896	3.355
9	1.383	1.833	2.262	2.821	3.250
10	1.372	1.812	2.228	2.764	3.169
∞	1.282	1.645	1.960	2.326	2.576

B.3 VALUES OF THE REJECTION QUOTIENT, Q_{CRIT}

Number of Observations	Q_{crit} 90% Confidence	96% Confidence	99% Confidence
3	0.94	0.98	0.99
4	0.76	0.85	0.93
5	0.64	0.73	0.82
6	0.56	0.64	0.74
7	0.51	0.59	0.68
8	0.47	0.54	0.63
9	0.44	0.51	0.60
10	0.41	0.48	0.57

B.4 VALUES OF THE ION SIZE PARAMETER, ALPHA

α (pm)	Ion
1100	Th^{4+}, Zr^{4+}, Ce^{4+}, Sn^{4+}
900	H^{+}, Al^{3+}, Fe^{3+}, Cr^{3+}, In^{3+}
800	Mg^{2+}, Be^{2+}
600	Li^{+}, Ca^{2+}, Cu^{2+}, Zn^{2+}, Sn^{2+}, Mn^{2+}, Fe^{2+}, Ni^{2+}, Co^{2+}
500	Sr^{2+}, Ba^{2+}, Cd^{2+}, Hg^{2+}, S2-, $Fe(CN)_6^{4-}$
450	Na^{+}, ClO_2^{-}, IO_3^{-}, HCO_3^{-}, $H_2PO_4^{-}$, Pb^{2+}, CO_3^{2-}, SO_3^{2-}
400	Hg_2^{2+}, SO_4^{2-}, $S_2O_3^{2-}$, CrO_4^{2-}, HPO_4^{2-}, PO_4^{3-}, $Fe(CN)_6^{3-}$
350	OH^{-}, F^{-}, SCN^{-}, OCN^{-}, HS^{-}, ClO_3^{-}, ClO_4^{-}, BrO_3^{-}, IO_4^{-}, MnO_4^{-}
300	K^{+}, Cl^{-}, Br^{-}, I^{-}, CN^{-}, NO_2^{-}, NO_3^{-}
250	Rb^{+}, Cs^{+}, NH_4^{+}, Tl^{+}, Ag^{+}

B.5 DISSOCIATION CONSTANTS OF COMMON ACIDS

Acid	Formula	pK_a
Acetic acid	CH_3COOH	4.76
Alanine	$HOOCCH(CH_3)NH_3^{+}$	2.35 9.87
Arsenic acid	H_3AsO_4	2.22 6.98 11.50
Arsenious acid	H_3AsO_3	9.29
Benzoic acid	C_6H_5COOH	4.21
Butanoic acid	C_3H_7COOH	4.98
Carbonic acid	H_2CO_3	6.35 10.33
Chloroacetic acid	$ClCH_2COOH$	2.87
Citric acid	$C_3H_5O(COOH)_3$	3.12 4.76 6.40
Cyanic acid	HOCN	3.66
Dichloroacetic acid	$Cl_2CHCOOH$	1.26
Formic acid	HCOOH	3.75

Glycine	$HOOCCH_2NH_3^+$	2.35 9.78
Hydrazoic acid	HN_3	4.72
Hydrocyanic acid	HCN	9.21
Hydrogen peroxide	HOOH	11.65
Hydrogen sulfide	H_2S	7.02 13.9
Hypobromous acid	HBrO	8.60
Hypochlorous acid	HClO	7.53
Hypoiodous acid	HIO	10.64
Maleic acid	HOOCCH=CHCOOH	1.92 6.22
Malonic acid	$HOOCCH_2COOH$	2.85 5.70
Nitrous acid	HNO_2	3.29
Oxalic acid	HOOCCOOH	1.27 4.27
Phenol	C_6H_5OH	9.98
Phosphoric acid	H_3PO_4	2.15 7.20 12.15
Propanoic acid	CH_3CH_2COOH	4.87
Succinic acid	$HOOCCH_2CH_2COOH$	4.21 5.64
Sulfamic acid	NH_3SO_3	1.0
Sulfuric acid (2nd)	HSO_4^-	1.99
Sulfurous acid	H_2SO_3	1.76 7.19
Tartaric acid	HOOCCHOHCHOHCOOH	3.04 4.37
Trichloroacetic acid	Cl_3CCOOH	0.66
Valine	$HOOCCH(NH_3^+)CH(CH_3)_2$	2.29 9.72

B.6 DISSOCIATION CONSTANTS OF COMMON BASES

Base	Formula	pK_b
Ammonia	NH_3	4.76
Aniline	$C_6H_5NH_2$	9.40
Cyclohexylamine	$C_6H_{11}NH_2$	3.36
Dimethylamine	$(CH_3)_2NH$	3.23
Ethylamine	$C_2H_5NH_2$	3.37
Ethylenediamine	$NH_2CH_2CH_2NH_2$	3.35 6.48
Hydrazine	H_2NNH_2	5.77
Hydroxylamine	$HONH_2$	8.04
Methylamine	CH_3NH_2	3.38
Piperidine	$C_5H_{10}NH$	2.88
Propylamine	$C_3H_7NH_2$	3.43
Pyridine	C_5H_5N	8.78
Pyrrolidine	C_4H_8NH	2.69
Triethanolamine	$(HOC_2H_4)_3N$	6.24
Trimethylamine	$(CH_3)_3N$	4.20

B.7 pK_a VALUES OF ACID-BASE INDICATORS

Indicator	pK_a	Colour change (acid/base)
Thymol blue	1.65	red/yellow
Quinaldine red	2.60	colourless/red
Methyl orange	3.39	red/orange
2,6-Dinitrophenol	3.67	colourless/yellow
Bromophenol blue	4.10	yellow/blue
Bromocresol green	4.90	yellow/blue
Methyl red	4.97	red/yellow
Chlorophenol red	6.22	yellow/red
Bromocresol purple	6.38	yellow/purple
p-Nitrophenol	6.95	colourless/yellow
Phenol red	7.96	yellow/red
m-Nitrophenol	8.30	colourless/yellow
Thymol blue	9.20	yellow/blue
Phenolphthalein	9.50	colourless/red
Thymolphthalein	10.0	colourless/blue

B.8 STANDARD POTENTIALS

Half Reaction	Standard Potential (V)
$Ag(I) + e^- \rightarrow Ag$	+0.799
$AgBr(s) + e^- \rightarrow Ag + Br^-$	+0.071
$AgCl(s) + e^- \rightarrow Ag + Cl^-$	+0.222
	+0.197 (satd KCl)
$AgI(s) + e^- \rightarrow Ag + I^-$	-0.152
$H_3AsO_4 + 2H^+ + 2e^- \rightarrow H_3AsO_3 + H_2O$	+0.559
$H_3AsO_3 + 3H^+ + 3e^- \rightarrow As + 3H_2O$	+0.248
$BiO^+ + 2H^+ + 3e^- \rightarrow Bi + H_2O$	+0.32
$BiOCl(s) + 2H^+ + 3e^- \rightarrow Bi + H_2O + Cl^-$	+0.160
$Br_2(aq) + 2e^- \rightarrow 2Br^-$	+1.087
$Br_3^- + 2e^- \rightarrow 3Br^-$	+1.051
$BrO^- + H_2O + 2e^- \rightarrow Br^- + 2OH^-$	+0.76
$2CO_2(g) + 2H^+ + 2e^- \rightarrow H_2C_2O_4$	-0.49
$Cd(II) + 2e^- \rightarrow Cd$	-0.402
$Ce(IV) + e^- \rightarrow Ce(III)$	+1.70 (1 F $HClO_4$)
	+1.60 (1 F HNO_3)
	+1.44 (1 F H_2SO_4)
	+1.28 (1 F HCl)
$Cl_2(aq) + 2e^- \rightarrow 2Cl^-$	+1.395
$HClO + H^+ + e^- \rightarrow 1/2Cl_2(g) + H_2O$	+1.63
$Co(II) + 2e^- \rightarrow Co$	-0.277
$Cr(II) + 2e^- \rightarrow Cr$	-0.91
$Cr(III) + e^- \rightarrow Cr(II)$	-0.41
$CrO_4^{2-} + 4H_2O + 3e^- \rightarrow Cr(OH)_3(s) + 5OH^-$	-0.13
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr(III) + 7H_2O$	+1.33
$Cu(II) + 2e^- \rightarrow Cu$	+0.337
$Cu(I) + e^- \rightarrow Cu$	+0.518
$Cu(II) + e^- \rightarrow Cu(I)$	+0.159
$F_2 + 2e^- \rightarrow 2F^-$	+2.87

$Fe(II) + 2e^- \rightarrow Fe$	-0.440
$Fe(III) + e^- \rightarrow Fe(II)$	+0.771
	+0.732 (1 F HCl)
	+0.68 (1 F H_2SO_4)
$2H^+ + 2e^- \rightarrow H_2$	+0.0000
$Hg^{2+} + 2e^- \rightarrow Hg$	+0.854
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	+0.792
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	+0.907
$Hg_2Br_2(s) + 2e^- \rightarrow 2Hg + 2Br^-$	+0.139
$Hg_2Cl_2(s) + 2e^- \rightarrow 2Hg + 2Cl^-$	+0.268
	+0.242 (satd KCl)
$Hg_2I_2(s) + 2e^- \rightarrow 2Hg + 2I^-$	-0.040
$Hg_2SO_4(s) + 2e^- \rightarrow 2Hg + SO_4^{2-}$	+0.614
$I_2(aq) + 2e^- \rightarrow 2I^-$	+0.622
$I_3^- + 2e^- \rightarrow 3I^-$	+0.536
$HOI + H^+ + 2e^- \rightarrow I^- + H_2O$	+0.99
$IO_3^- + 6H^+ + 5e^- \rightarrow 1/2I_2 + 3H_2O$	+1.178
$Mn(II) + 2e^- \rightarrow Mn$	-1.182
$Mn(III) + e^- \rightarrow Mn(II)$	+1.542
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	+0.558
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2(s) + 2H_2O$	+1.69
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn(II) + 4H_2O$	+1.51
$MoO_4^{2-} + 4H_2O + 6e^- \rightarrow Mo + 8OH^-$	-1.05
$N_2(g) + 2H_2O + 4H^+ + 2e^- \rightarrow 2NH_3OH^+$	-1.87
$N_2H_5^+ + 3H^+ + 2e^- \rightarrow 2NH_4^+$	+1.275
$NO_3^- + 3H^+ + 2e^- \rightarrow HNO_2 + H_2O$	+0.94
$NO_3^- + 4H^+ + 3e^- \rightarrow NO(g) + 2H_2O$	+0.96
$HNO_2 + H^+ + e^- \rightarrow NO(g) + H_2O$	+1.00
$Ni(II) + 2e^- \rightarrow Ni$	-0.231
$O_2(g) + 2H^+ + 2e^- \rightarrow H_2O_2$	+0.682
$O_2(g) + 4H^+ + 4e^- \rightarrow H_2O$	+1.229
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	+1.776
$OsO_4(s) + 8H^+ + 8e^- \rightarrow Os + 4H_2O$	+0.85

$Pb(II) + 2e^- \rightarrow Pb$	-0.126
$PbSO_4(s) + 2e^- \rightarrow Pb + SO_4^{2-}$	-0.355
$PbO_2(s) + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4(s) + 2H_2O$	+1.685
$S(s) + 2H^+ + 2e^- \rightarrow H_2S(aq)$	+0.141
$SO_4^{2-} + 4H^+ + 2e^- \rightarrow SO_2(g) + 2H_2O$	+0.17
$SO_4^{2-} + H_2O + 2e^- \rightarrow SO_3^{2-} + 2OH^-$	-0.93
$S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$	+2.01
$S_4O_6^{2-} + 2e^- \rightarrow 2S_2O_3^{2-}$	+0.090
$SbO_3^- + H_2O + 2e^- \rightarrow SbO_2^- + 2OH^-$	-0.589 (10 F NaOH)
$H_3SbO_4 + 2H^+ + 2e^- \rightarrow H_3SbO_3 + H_2O$	+0.75
$SbO_2^- + 2H_2O + 3e^- \rightarrow Sb + 4OH^-$	+0.675 (10 F NaOH)
$SbO^+ + 2H^+ + 3e^- \rightarrow Sb + H_2O$	+0.212
$Sn(II) + 2e^- \rightarrow Sn$	-0.136
$Sn(IV) + 2e^- \rightarrow Sn(II)$	+0.154
	+0.140 (1 F HCl)
$Ti(II) + 2e^- \rightarrow Ti$	-1.630
$Ti(III) + e^- \rightarrow Ti(II)$	-0.368
$TiO^{2+} + 2H^+ + e^- \rightarrow Ti(III) + H_2O$	+0.100
$Tl(I) + e^- \rightarrow Tl$	-0.336
$Tl(III) + 2e^- \rightarrow Tl(I)$	+1.28
$TlCl(s) + e^- \rightarrow Tl + Cl^-$	-0.557
$U(III) + 3e^- \rightarrow U$	-1.80
$U(IV) + e^- \rightarrow U(III)$	-0.61
$UO_2^{2+} + 4H^+ + 2e^- \rightarrow U(IV) + 2H_2O$	+0.334
$V(II) + 2e^- \rightarrow V$	-1.18
$V(III) + e^- \rightarrow V(II)$	-0.255
$VO^{2+} + 2H^+ + e^- \rightarrow V(III) + H_2O$	+0.337
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	+1.000
$Zn(II) + 2e^- \rightarrow Zn$	-0.764

B.9 STANDARD POTENTIALS OF REDOX INDICATORS

Indicator	E°	Colour Change (Ox/Red)
5-Nitroferroin	+1.25	blue/purple
Ferroin	+1.11	blue/red
5-Methylferroin	+1.02	Blue-green/red
Erioglaucine A	+0.98	Blue-red/yellow-green
Diphenylaminesulfonic acid	+0.84	purple/colourless
Diphenlyamine	+0.76	violet/colourless
Methylene blue	+0.53	blue/colourless
Indigotetrasulfonate	+0.36	blue/colourless
Phenosafranine	+0.28	red/colourless
m-Bromophenolindophenol	+0.25	red/colourless