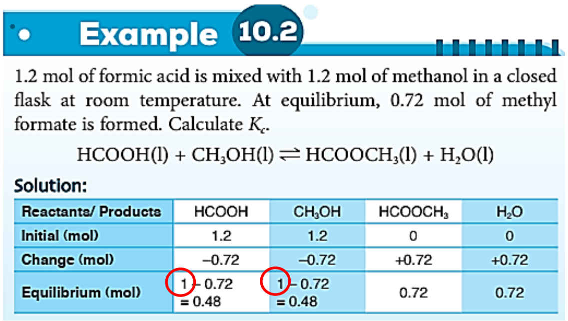
Title: **STRATEGIC STPM PRE-U TEXT CHEMISTRY SEMESTER 2 (2025)**
Book Code: **AAEVSC2675031A2**
Author: **Peh Kok Hua, Lee Sze Yien,Chong Boon Weng, Peter Yip**

<table>
<tr><th>Page number</th><th>Section</th><th>Error</th><th>Correction</th></tr>
<tr>
<td>38</td>
<td>10.1</td>
<td>

Example 10.2
1.2 mol of formic acid is mixed with 1.2 mol of methanol in a closed flask at room temperature. At equilibrium, 0.72 mol of methyl formate is formed. Calculate K_c.
$HCOOH(l) + CH_3OH(l) \rightleftharpoons HCOOCH_3(l) + H_2O(l)$
Solution:
<table>
<tr><th>Reactants/ Products</th><th>HCOOH</th><th>CH_3OH</th><th>$HCOOCH_3$</th><th>H_2O</th></tr>
<tr><td>Initial (mol)</td><td>1.2</td><td>1.2</td><td>0</td><td>0</td></tr>
<tr><td>Change (mol)</td><td>–0.72</td><td>–0.72</td><td>+0.72</td><td>+0.72</td></tr>
<tr><td>Equilibrium (mol)</td><td>1 – 0.72
= 0.48</td><td>1 – 0.72
= 0.48</td><td>0.72</td><td>0.72</td></tr>
</table>
</td>
<td>1.2 – 0.72
= 0.48</td>
</tr>
<tr>
<td>51</td>
<td>10.1</td>
<td>
1. The synthesis of hydrogen iodide is a reversible reaction between hydrogen gas and iodine gas.

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ ΔH = positive

This reaction is endothermic in the forward direction.

2. The reaction takes place in the gas phase, and no change in the number of moles of gas occurs during the reaction.

3. Temperature:
(a) Since the forward reaction is endothermic, higher temperature favours the formation of HI.
(b) However, excessively high temperatures may cause decomposition of HI
</td>
<td>
1. The synthesis of hydrogen iodide is a reversible reaction between hydrogen gas and iodine gas.

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ ΔH = -9 kJ mol^{-1}

This reaction is exothermic in the forward direction.

2. The reaction takes place in the gas phase, and no change in the number of moles of gas occurs during the reaction.

3. Temperature:
(a) Since the forward reaction is exothermic, lower temperature favours the formation of HI.(b) However, too low temperatures may slow down the reaction rate.
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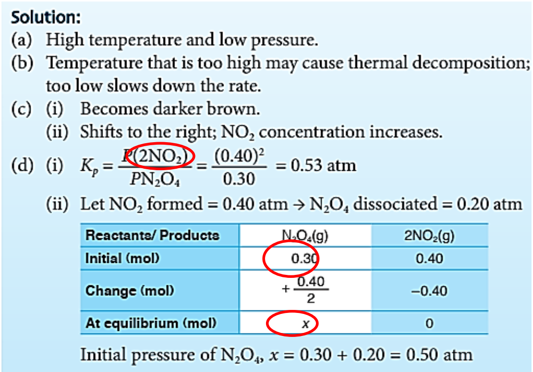
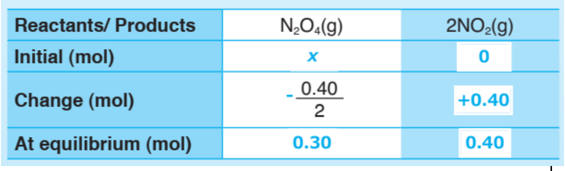
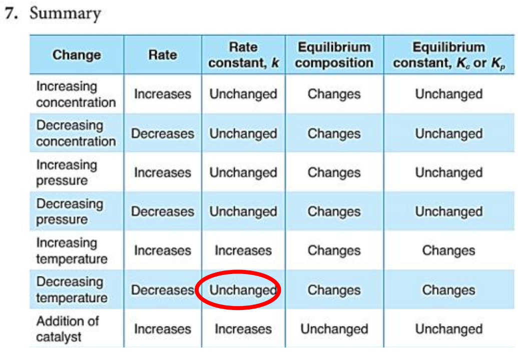
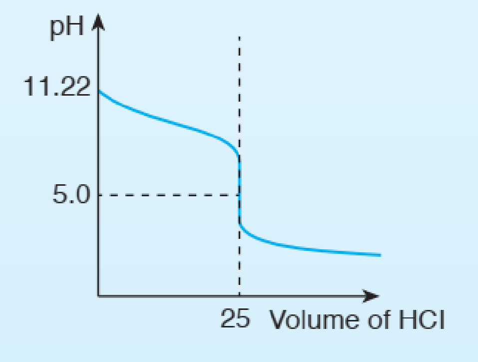
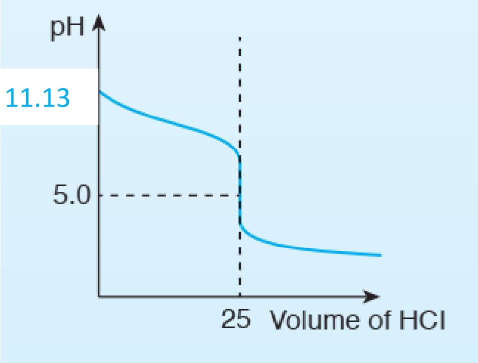
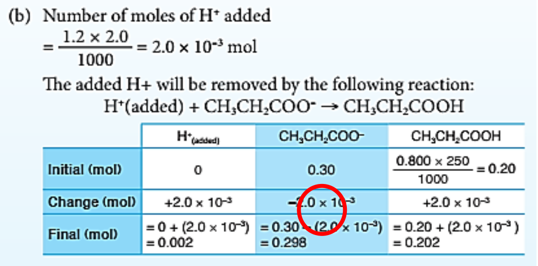
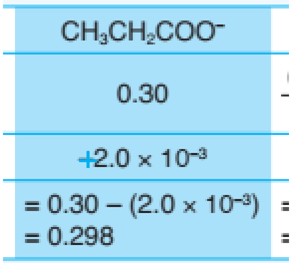
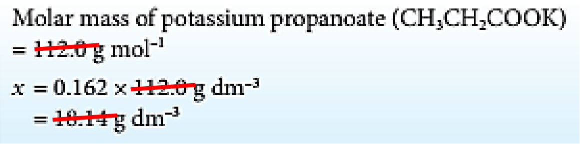
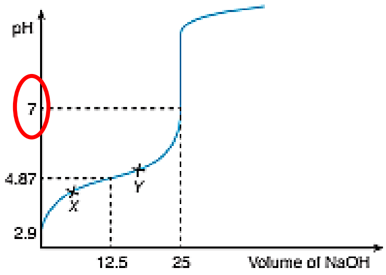
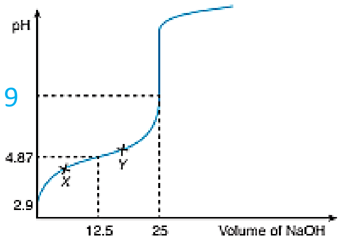
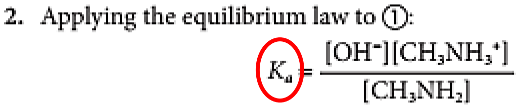
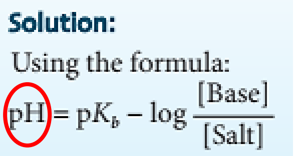
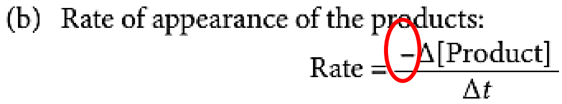
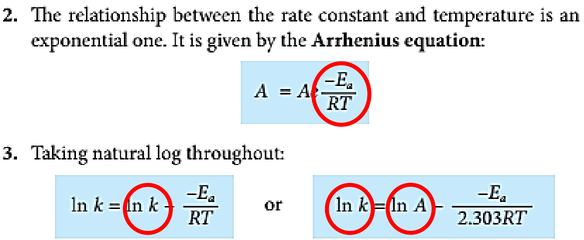
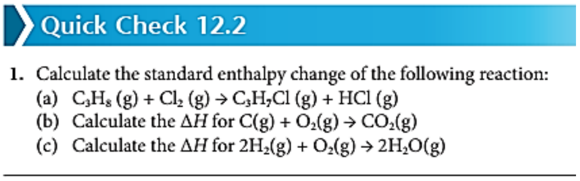
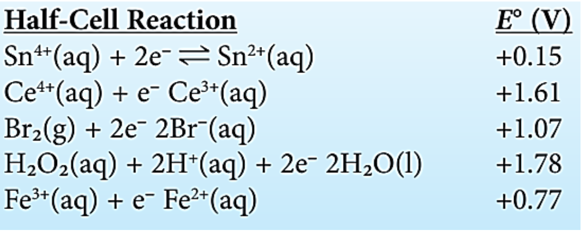
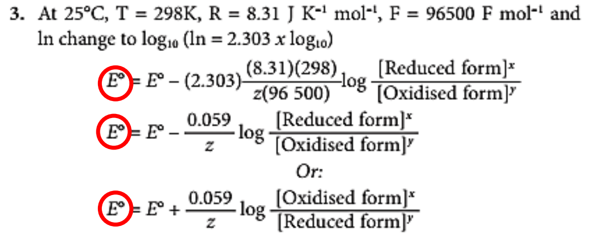
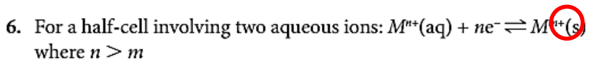
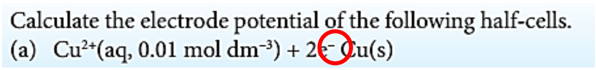
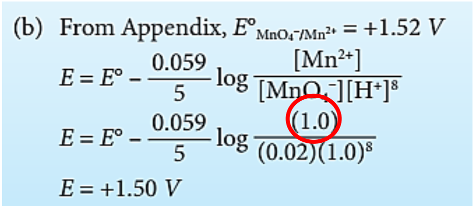
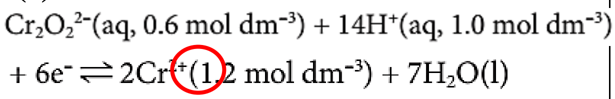
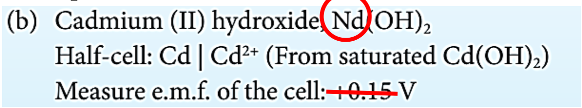
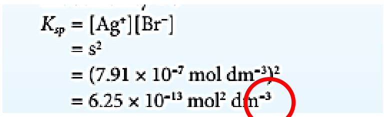
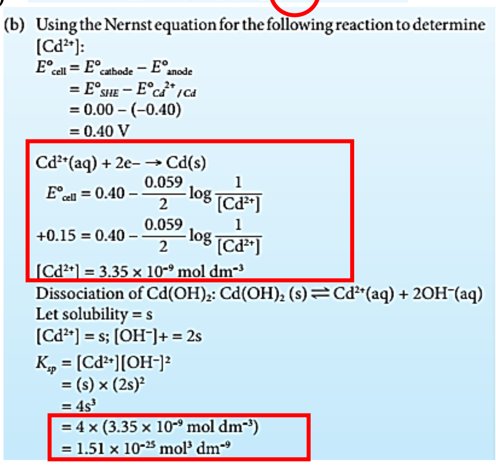
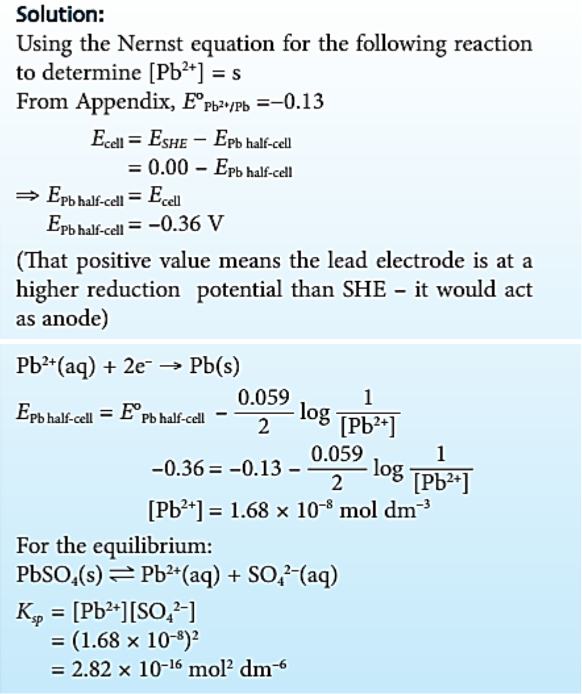
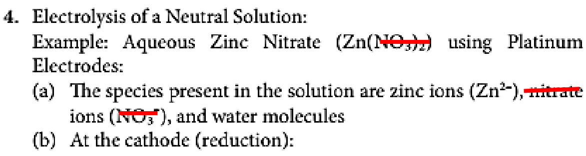
<tr>
<td>57</td>
<td>10.1</td>
<td>
Example 10.16

Solution:
(a) High temperature and low pressure.
(b) Temperature that is too high may cause thermal decomposition; too low slows down the rate.
(c) (i) Becomes darker brown.
(ii) Shifts to the right; NO_2 concentration increases.
(d) (i) $K_p = \frac{P(2NO_2)}{PN_2O_4} = \frac{(0.40)^2}{0.30} = 0.53$ atm
(ii) Let NO_2 formed = 0.40 atm → N_2O_4 dissociated = 0.20 atm
<table>
<tr><th>Reactants/ Products</th><th>$N_2O_4(g)$</th><th>$2NO_2(g)$</th></tr>
<tr><td>Initial (mol)</td><td>0.30</td><td>0.40</td></tr>
<tr><td>Change (mol)</td><td>$+\frac{0.40}{2}$</td><td>–0.40</td></tr>
<tr><td>At equilibrium (mol)</td><td>x</td><td>0</td></tr>
</table>
Initial pressure of N_2O_4, $x = 0.30 + 0.20 = 0.50$ atm
</td>
<td>
d (i) $(P_{NO_2})^2$
(ii)

<table>
<tr><th>Reactants/ Products</th><th>$N_2O_4(g)$</th><th>$2NO_2(g)$</th></tr>
<tr><td>Initial (mol)</td><td>x</td><td>0</td></tr>
<tr><td>Change (mol)</td><td>$-\frac{0.40}{2}$</td><td>+0.40</td></tr>
<tr><td>At equilibrium (mol)</td><td>0.30</td><td>0.40</td></tr>
</table>
</td>
</tr>
</table>

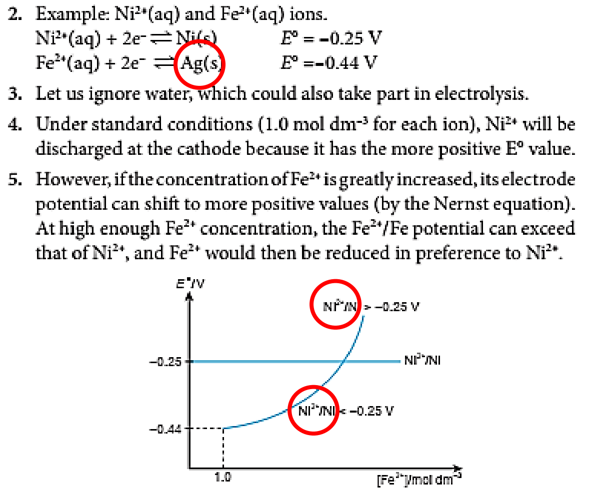
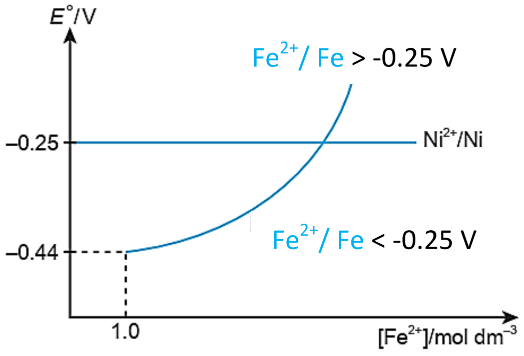
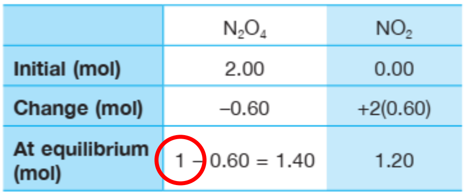
<table>
<tr><td>59</td><td>10.1</td><td>
7. Summary
<table>
<tr><th>Change</th><th>Rate</th><th>Rate constant, k</th><th>Equilibrium composition</th><th>Equilibrium constant, K_c or K_p</th></tr>
<tr><td>Increasing concentration</td><td>Increases</td><td>Unchanged</td><td>Changes</td><td>Unchanged</td></tr>
<tr><td>Decreasing concentration</td><td>Decreases</td><td>Unchanged</td><td>Changes</td><td>Unchanged</td></tr>
<tr><td>Increasing pressure</td><td>Increases</td><td>Unchanged</td><td>Changes</td><td>Unchanged</td></tr>
<tr><td>Decreasing pressure</td><td>Decreases</td><td>Unchanged</td><td>Changes</td><td>Unchanged</td></tr>
<tr><td>Increasing temperature</td><td>Increases</td><td>Increases</td><td>Changes</td><td>Changes</td></tr>
<tr><td>Decreasing temperature</td><td>Decreases</td><td>Unchanged</td><td>Changes</td><td>Changes</td></tr>
<tr><td>Addition of catalyst</td><td>Increases</td><td>Increases</td><td>Unchanged</td><td>Unchanged</td></tr>
</table></td><td>Decreases</td></tr>
<tr><td>65</td><td>10.2</td><td>3. Example: Dissolution of HNO_3 in water: When hydrogen nitrate dissolves in water to form hydrochloric acid, the reaction is:</td><td>3. Example: Dissolution of HNO_3 in water: When hydrogen nitrate dissolves in water to form hydronium ion, the reaction is:</td></tr>
<tr><td>79</td><td>10.2</td><td>4. Example : Sulphurous acid dissociates in two stages:</td><td>4. Example : Carbonic acid dissociates in two stages:</td></tr>
<tr><td>87</td><td>10.2</td><td>Example 10.30
Solution:
(b) For weak base:
$\alpha = \frac{[OH^-]^{2}}{c} = \frac{3.16 \times 10^{-3}}{0.10} = 0.0316$
$\alpha = 0.0316$ or 3.16%</td><td>$\alpha = \frac{[OH^-]}{c} = \frac{3.16 x 10^{-3}}{0.10} = 0.0316$</td></tr>
<tr><td>88</td><td>10.2</td><td>Example 10.33
Solution:
Using the formula: $[H_3O^+]$</td><td>$[OH^-]$</td></tr>
<tr><td>100</td><td>10.2</td><td>
</td><td>
</td></tr>
<tr><td>104</td><td>10.2</td><td>5. Final Henderson – Hasselbalch equation for acid:
$\therefore \ pH = K_a - \log\frac{[Acid]}{[Salt]}$</td><td>$pK_a$</td></tr>
<tr><td>106</td><td>10.2</td><td>Example 10.43
Solution:
Using the formula:
$pH = pK_b - \log\frac{[Base]}{[Salt]}$</td><td>$pK_a$</td></tr>
</table>

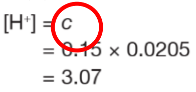
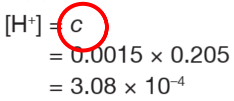
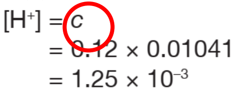
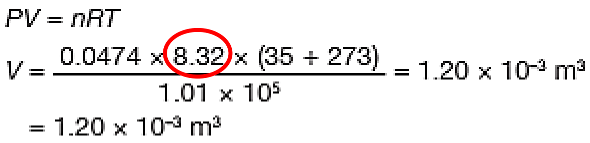
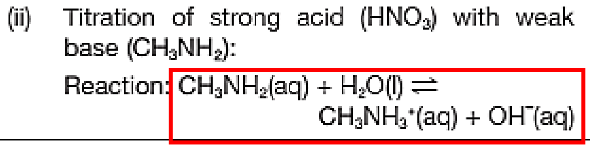
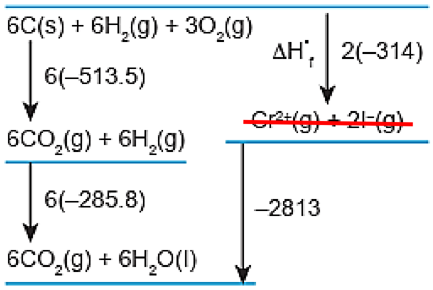
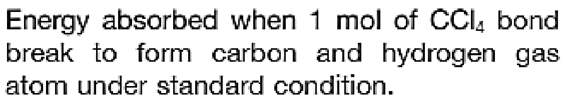
<table>
<tr><td>107</td><td>10.2</td><td>(b) Number of moles of H^+ added
$= \frac{1.2 \times 2.0}{1000} = 2.0 \times 10^{-3}$ mol
The added H+ will be removed by the following reaction:
$H^+(added) + CH_3CH_2COO^- \rightarrow CH_3CH_2COOH$
<table>
<tr><td></td><td>$H^+_{(added)}$</td><td>$CH_3CH_2COO^-$</td><td>CH_3CH_2COOH</td></tr>
<tr><td>Initial (mol)</td><td>0</td><td>0.30</td><td>$\frac{0.800 \times 250}{1000} = 0.20$</td></tr>
<tr><td>Change (mol)</td><td>$+2.0 \times 10^{-3}$</td><td>-2.0×10^{-3}</td><td>$+2.0 \times 10^{-3}$</td></tr>
<tr><td>Final (mol)</td><td>$= 0 + (2.0 \times 10^{-3})$
$= 0.002$</td><td>$= 0.30 - (2.0 \times 10^{-3})$
$= 0.298$</td><td>$= 0.20 + (2.0 \times 10^{-3})$
$= 0.202$</td></tr>
</table></td><td>$CH_3CH_2COO^-$
0.30
$+2.0 \times 10^{-3}$
$= 0.30 - (2.0 \times 10^{-3})$
$= 0.298$</td></tr>
<tr><td>110</td><td>10.2</td><td>Example 10.44
Molar mass of potassium propanoate (CH_3CH_2COOK)
= ~~112.0~~ g mol^{-1}
x = 0.162 × ~~112.0~~ g dm^{-3}
= ~~18.14~~ g dm^{-3}</td><td>Molar mass of potassium propanoate (CH_3CH_2COOK) = 112.1 g mol^{-1}
X = 0.162 x 112.1 g dm^{-3}
= 18.16 g dm^{-3}</td></tr>
<tr><td>110</td><td>10.2</td><td>2. The pH graph is as follows:
 </td><td>2. The pH graph is as follows:
 </td></tr>
<tr><td>105</td><td>10.2</td><td>2. Applying the equilibrium law to ①:
$K_a = \frac{[OH^-][CH_3NH_3^+]}{[CH_3NH_2]}$</td><td>$K_b = \frac{[OH^-][CH_3NH_3^+]}{[CH_3NH_2]}$</td></tr>
<tr><td>106</td><td>10.2</td><td>Example 10.42
Solution:
Using the formula:
$pH = pK_b - \log \frac{[Base]}{[Salt]}$</td><td>Using the formula:
$pOH = pK_b - \log \frac{[Base]}{[Salt]}$</td></tr>
<tr><td>144</td><td>11.1</td><td>3.
(b) Rate of appearance of the products:
$Rate = \frac{-\Delta[Product]}{\Delta t}$</td><td>$Rate = \frac{\Delta\ [Product]}{\Delta t}$</td></tr>
<tr><td>157</td><td>11.2</td><td>2. The relationship between the rate constant and temperature is an exponential one. It is given by the Arrhenius equation:
$A = Ae^{\frac{-E_a}{RT}}$
3. Taking natural log throughout:
$\ln k = \ln k - \frac{-E_a}{RT}$ or $\ln k = \ln A - \frac{-E_a}{2.303RT}$</td><td>2. $k = Ae^{\frac{-Ea}{RT}}$
3. $\ln k = \ln A - \frac{Ea}{RT}$
or
$\log k = \log A - \frac{Ea}{2.303RT}$</td></tr>
<tr><td>190</td><td>STPM Practice 11</td><td>1. (d)
What is the order of reaction with respect to hydrogen peroxide?</td><td>What is the order of reaction with respect to ester/ ethyl ethanoate?</td></tr>
</table>

200	**12.1**	**1.** The change in enthalpy that occurs when **one mole of a compound is converted into gaseous atoms**. All bonds in the compound are broken in atomization and none are formed, so enthalpies of atomization are always positive.	The standard enthalpy change of atomisation of an element is the heat absorbed when one mole of free gaseous atoms is formed its element under standard condition. All bonds in the compound are broken in atomization and none are formed, so enthalpies of atomization are always positive.
201	**12.1**	Example 12.31 CH_4(g) + Br_2(l) → CH_3Br(g) + HBr(g) Given the enthalpy of formation of CH_4 is -74.8 kJmol-1.While enthalpies of formation of CH_3Br and HBr are –37.2 kJ mo$_{-1}$ and –36.4 kJ mol$_{-1}$ respectively. Calculate the enthalpy change for the reaction. (**change question to new one**)	Using the average bond enthalpies in table above, calculate the enthalpy change for the combustion of methane. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
205	**12.1**	3……………… If lattice energy lower than enthalpy of hydration, the enthalpy of solution has an endothermic value. The salt involve often insoluble. 5. ………………If lattice energy greater than enthalpy of hydration, the enthalpy of solution has an exothermic value. The salt involve often soluble.	……………… If lattice energy lower than enthalpy of hydration, the enthalpy of solution has an exothermic value. The salt involve often soluble. ………………If lattice energy greater than enthalpy of hydration, the enthalpy of solution has an endothermic value. The salt involve often insoluble.
206	**12.1**	Quick Check 12.2  Quick Check 12.2 1. Calculate the standard enthalpy change of the following reaction: (a) C_3H_8 (g) + Cl_2 (g) → C_3H_7Cl (g) + HCl (g) (b) Calculate the Δ*H* for C(g) + O_2(g) → CO_2(g) (c) Calculate the Δ*H* for $2H_2$(g) + O_2(g) → $2H_2O$(g) ΔH List missing	(refer data of average bond energies at page 201)
217	**13.1**	**9.** $Fe^{3+} + e^{-} \rightarrow 2e^{2+}$ $Cl_2 + 2e^{-} \rightarrow 2Cl_2$	$Fe^{3+} + e^{-} \rightarrow Fe^{2+}$ $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$
238	**13.2**	Example 13.10  **Half-Cell Reaction** — ***E*° (V)** $Sn^{4+}(aq) + 2e^{-} \rightleftharpoons Sn^{2+}(aq)$ +0.15 $Ce^{4+}(aq) + e^{-}$ $Ce^{3+}(aq)$ +1.61 $Br_2(g) + 2e^{-}$ $2Br^{-}(aq)$ +1.07 $H_2O_2(aq) + 2H^{+}(aq) + 2e^{-}$ $2H_2O(l)$ +1.78 $Fe^{3+}(aq) + e^{-}$ $Fe^{2+}(aq)$ +0.77	**Half-Cell Reaction** ***E* °(V)** $Sn^{4+}(aq) + 2e- \rightleftharpoons Sn^{2+}(aq)$ +0.15 $Ce^{4+}(aq) + e- \rightleftharpoons Ce^{3+}(aq)$ +1.61 $Br_2(g) + 2e- \rightleftharpoons 2Br^{-}(aq)$ +1.07 $H_2O_2(aq) + 2H^{+}(aq) + 2e- \rightleftharpoons 2H_2O(l)$ +1.78 $Fe^{3+}(aq) + e- \rightleftharpoons$ Fe2+(aq) +0.77
242	**13.2**	Example 13.12 $E°_{cell}$ of the overall reaction is negative. Thus, $Cr_2O_7^{2-}$ does not react with Cl^{-} spontaneously under standard conditions.	$E°_{cell}$ of the overall reaction is positive. Thus, MnO_4^{-} react with Fe^{2+} spontaneously under standard conditions.

245	**13.3**	 3. At 25°C, T = 298K, R = 8.31 J K⁻¹ mol⁻¹, F = 96500 F mol⁻¹ and ln change to $\log_{10}$ (ln = 2.303 *x* $\log_{10}$) $E° = E° - (2.303)\frac{(8.31)(298)}{z(96\ 500)}\log\frac{[\text{Reduced form}]^x}{[\text{Oxidised form}]^y}$ $E° = E° - \frac{0.059}{z}\log\frac{[\text{Reduced form}]^x}{[\text{Oxidised form}]^y}$ *Or:* $E° = E° + \frac{0.059}{z}\log\frac{[\text{Oxidised form}]^x}{[\text{Reduced form}]^y}$	$E = E° - (2.303)\frac{(8.31)(298)}{z(96\ 500)}\log\frac{[\text{Reduced form}]^x}{[\text{Oxidised form}]^y}$ $E = E° - \frac{0.059}{z}\log\frac{[\text{Reduced form}]^x}{[\text{Oxidised form}]^y}$ *Or:* $E = E° + \frac{0.059}{z}\log\frac{[\text{Oxidised form}]^x}{[\text{Reduced form}]^y}$
246	**13.3**	 6. For a half-cell involving two aqueous ions: $M^{n+}(aq) + ne^- \rightleftharpoons M^{m+}(s)$ where $n > m$	$M^{n+}(aq) + ne^-$ L $M^{m+}(aq)$
246	**13.3**	Example 13.14  Calculate the electrode potential of the following half-cells. (a) $Cu^{2+}(aq, 0.01\ mol\ dm^{-3}) + 2e^-\ Cu(s)$	$Cu^{2+}(aq, 0.01\ mol\ dm^{-3}) + 2e^- \rightleftharpoons Cu(s)$
247	**13.3**	Example 13.14 Solution:  (b) From Appendix, $E°_{MnO_4^-/Mn^{2+}} = +1.52\ V$ $E = E° - \frac{0.059}{5}\log\frac{[Mn^{2+}]}{[MnO_4^-][H^+]^8}$ $E = E° - \frac{0.059}{5}\log\frac{(1.0)}{(0.02)(1.0)^8}$ $E = +1.50\ V$	$E = E° - \frac{0.059}{5} log \frac{(0.1)}{(0.02)(1.0)^8}$
247	**13.3**	Quick Check 13.4 4(d)  $Cr_2O_2^{2-}(aq, 0.6\ mol\ dm^{-3}) + 14H^+(aq, 1.0\ mol\ dm^{-3}) + 6e^- \rightleftharpoons 2Cr^{3+}(1.2\ mol\ dm^{-3}) + 7H_2O(l)$	Cr_2O_{22}–$(aq, 0.6\ mol\ dm^{-3})$ + $14H^+(aq, 1.0\ mol\ dm^{-3})$ + 6e– ⇌ $2Cr^{2+}(aq, 1.2\ mol\ dm^{-3}) + 7H_2O(l)$
258	**13.3**	Example 13.22  (b) Cadmium (II) hydroxide, $Nd(OH)_2$ Half-cell: Cd | Cd^{2+} (From saturated $Cd(OH)_2$) Measure e.m.f. of the cell: ~~+0.15~~ V	(b) Cadmium (II) hydroxide, $Cd(OH)_2$ Half-cell: Cd | Cd^{2+} (From saturated $Cd(OH)_2$) Measure e.m.f. of the cell: +0.54 V
258	**13.3**	Example 13.22 Solution (a)  $K_{sp} = [Ag^+][Br^-]$ $= s^2$ $= (7.91 \times 10^{-7}\ mol\ dm^{-3})^2$ $= 6.25 \times 10^{-13}\ mol^2\ dm^{-3}$  (b) Using the Nernst equation for the following reaction to determine $[Cd^{2+}]$: $E°_{cell} = E°_{cathode} - E°_{anode}$ $= E°_{SHE} - E°_{Cd^{2+}/Cd}$ $= 0.00 - (-0.40)$ $= 0.40\ V$ $Cd^{2+}(aq) + 2e- \rightarrow Cd(s)$ $E°_{cell} = 0.40 - \frac{0.059}{2}\log\frac{1}{[Cd^{2+}]}$ $+0.15 = 0.40 - \frac{0.059}{2}\log\frac{1}{[Cd^{2+}]}$ $[Cd^{2+}] = 3.35 \times 10^{-9}\ mol\ dm^{-3}$ Dissociation of $Cd(OH)_2$: $Cd(OH)_2\ (s) \rightleftharpoons Cd^{2+}(aq) + 2OH^-(aq)$ Let solubility = s $[Cd^{2+}] = s; [OH^-]+ = 2s$ $K_{sp} = [Cd^{2+}][OH^-]^2$ $= (s) \times (2s)^2$ $= 4s^3$ $= 4 \times (3.35 \times 10^{-9}\ mol\ dm^{-3})$ $= 1.51 \times 10^{-25}\ mol^3\ dm^{-9}$	$K_{sp} = 6.25 \times 10^{-13}\ mol^2\ dm^{-6}$ (a) Since Cd has a more negative reduction potential, Cd acts as anode (oxidation): $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$ $E_{cell} = 0.40 - \frac{0.059}{2} log \frac{[Cd^{2+}]}{1}$ $+0.54 = 0.40 - \frac{0.059}{2} log[Cd^{2+}]$ $[Cd^{2+}] = 1.78 \times 10^{-5}\ mol\ dm^{-3}$ Dissociation of $Cd(OH)_2$: $Cd(OH)_2\ (s) \rightleftharpoons Cd^{2+}(aq) + 2OH^-(aq)$ Let solubility = s $[Cd^{2+}] = s; [OH^-] = 2s$ $K_{sp} = [Cd^{2+}]\ [OH^-]^2$ $= (s)\ x\ (2s)^2$ $= 4s^3$ $= 4\ x\ (1.78 \times 10^{-5}\ mol\ dm^{-3})^3$ $= 2.25 \times 10^{-14}\ mol^3\ dm^{-9}$

260	**13.3**	Example 13.23  **Solution:** Using the Nernst equation for the following reaction to determine $[Pb^{2+}] = s$ From Appendix, $E^\circ_{Pb^{2+}/Pb} = -0.13$ $E_{cell} = E_{SHE} - E_{Pb\ half\text{-}cell}$ $= 0.00 - E_{Pb\ half\text{-}cell}$ $\Rightarrow E_{Pb\ half\text{-}cell} = E_{cell}$ $E_{Pb\ half\text{-}cell} = -0.36$ V (That positive value means the lead electrode is at a higher reduction potential than SHE – it would act as anode) $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ $E_{Pb\ half\text{-}cell} = E^\circ_{Pb\ half\text{-}cell} - \frac{0.059}{2}\log\frac{1}{[Pb^{2+}]}$ $-0.36 = -0.13 - \frac{0.059}{2}\log\frac{1}{[Pb^{2+}]}$ $[Pb^{2+}] = 1.68 \times 10^{-8}$ mol dm^{-3} For the equilibrium: $PbSO_4(s) \rightleftharpoons Pb^{2+}(aq) + SO_4^{2-}(aq)$ $K_{sp} = [Pb^{2+}][SO_4^{2-}]$ $= (1.68 \times 10^{-8})^2$ $= 2.82 \times 10^{-16}$ mol^2 dm^{-6}	From appendix: $E^\circ_{Pb^{2+}/Pb} = -0.13$ V Since Pb has a more negative reduction potential, Pb acts as anode (oxidation): $Pb \rightarrow Pb^{2+} + 2e^-$ $E^\circ_{cell} = 0 - (-0.13) = 0.13$ V Using the Nernst equation: $E_{cell} = E^\circ_{cell} - \frac{0.059}{2}\log\frac{[Pb^{2+}]}{1}$ Substitute $E_{cell} = 0.36$: $0.36 = 0.13 - \frac{0.059}{2}\log[Pb^{2+}]$ $[Pb^{2+}] = 1.60 \times 10^{-8}\ moldm^{-3}$ For the equilibrium: $PbSO_4(s) \rightleftharpoons Pb^{2+} + SO_4^{2-}$ If solubility $= s$, $[Pb^{2+}] = [SO_4^{2-}] = s$ Thus, $K_{sp} = s^2$ $K_{sp} = (1.60 \times 10^{-8})^2$ $K_{sp} = 2.6 \times 10^{-16}\ mol^2 dm^{-6}$
260	**13.3**	Quick Check 13.8 **3.** A copper electrode is immersed in saturated CuBr (aq) and connected to the SHE.The measured cell potential at 25°C is -0.22 V. Calculate the solubility product (*Ksp*) of CuBr	A copper electrode is immersed in saturated CuBr (aq) and connected to the SHE.The measured cell potential at 25°C is +0.28V. Calculate the solubility product (*Ksp*) of CuBr
265	**13.4**	 4. Electrolysis of a Neutral Solution: Example: Aqueous Zinc Nitrate (Zn(~~$NO_3)_2$~~) using Platinum Electrodes: (a) The species present in the solution are zinc ions (Zn^{2-}), ~~nitrate~~ ions (~~NO_3^-~~), and water molecules (b) At the cathode (reduction):	4. Electrolysis of a Neutral Solution: Example: Aqueous Zinc Nitrate (Zn(Br)$_2$) using Platinum Electrodes: (a) The species present in the solution are zinc ions (Zn^{2-})’ bromide ions (Br^-), and water molecules.

267	**13.4**	 2. Example: Ni^{2+}(aq) and Fe^{2+}(aq) ions. $Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$ $E° = -0.25$ V $Fe^{2+}(aq) + 2e^- \rightleftharpoons Ag(s)$ $E° = -0.44$ V 3. Let us ignore water, which could also take part in electrolysis. 4. Under standard conditions (1.0 mol dm^{-3} for each ion), Ni^{2+} will be discharged at the cathode because it has the more positive E° value. 5. However, if the concentration of Fe^{2+} is greatly increased, its electrode potential can shift to more positive values (by the Nernst equation). At high enough Fe^{2+} concentration, the Fe^{2+}/Fe potential can exceed that of Ni^{2+}, and Fe^{2+} would then be reduced in preference to Ni^{2+}. 	Fe^{2+} (aq) + $2e^-$ $\rightleftharpoons$ Fe (s) 
280	**13.4**	Quick Check 13.10 3. When a current of 1.5 A flows for 43 minutes through an aqueous solution containing gold ion, Aun+, 2.63 g of gold was deposited on the cathode. Calculate the charge on the gold ion.	3. When a current of 1.5 A flows for 43 minutes through an aqueous solution containing gold ion, Au^{n+}, 2.63 g of gold was deposited on the cathode. Calculate the charge on the gold ion.
288	**STPM Practice 13**	**Objective Questions** **1.** Three standard electrochemical cells are measured: $Mg/Mg^{2+} \| Cu^{2+}/Cu$ $E°_{cell}$ = ~~2.10 V~~ $Mg/Mg^{2+} \| H^+/H_2$ (Pt) $E°_{cell}$ = ~~1.56 V~~ $Cu^{2+}/Cu \| Fe^{3+}/Fe^{2+}$ (Pt) $E°_{cell}$ = ~~0.79 V~~	**1.** Three standard electrochemical cells are measured: $Mg/Mg^{2+} \| Cu^{2+}/Cu$ $E°_{cell}$ = 2.71V $Mg/Mg^{2+} \| H^+/H_2$ (Pt) $E°_{cell}$ = 2.37V $Cu^{2+}/Cu \| Fe^{3+}/Fe^{2+}$ (Pt) $E°_{cell}$ = 0.43V
292	**STPM Practice 13**	Structured and Essay Questions 3. . ………The measured cell potential is +0.42 V at 25 °C. Standard electrode potentials: $E°_{Ag^+}/Ag = +0.80\ V$, ~~$E°_{AgCl/Ag}$~~ ~~= 0.504~~ V, $E°_{Cu^{2+}/Cu} = +0.34\ V$, $E°_{Cu^{2+}/Cu^+}$ $= +0.15\ V$, $E°_{Cu^+/Cu} = +0.52\ V$	3. ………The measured cell potential is +0.26 V at 25 °C. Standard electrode potentials: $E°_{Ag^+}/Ag = +0.80\ V$, $E°_{Cu^{2+}/Cu} = +0.34\ V$, $E°_{Cu^{2+}/Cu^+}$ $= +0.15\ V$, $E°_{Cu^+/Cu} = +0.52\ V$
293	**STPM Practice 13**	7 (a) Identify the metal species which reacts with solid zinc to produce the minimum value of standard cell potential. Write the cell notation and calculate the of the reaction.	Identify the metal species which reacts with solid zinc to produce the minimum value of standard cell potential. Write the cell notation and calculate the $E°_{cell}$ of the reaction.
301	**Answers**	Chapter 10 Quick check 10.1 3 (a)  | N_2O_4 | NO_2 **Initial (mol)** | 2.00 | 0.00 **Change (mol)** | −0.60 | +2(0.60) **At equilibrium (mol)** | 1 – 0.60 = 1.40 | 1.20	2 – 0.60 = 1.40

304	**Answers**	Chapter 10 Quick check 10.9 3 (a)  $[H^+] = c$ $= 0.15 \times 0.0205$ $= 3.07$ 3 (b)  $[H^+] = c$ $= 0.0015 \times 0.205$ $= 3.08 \times 10^{-4}$	$[H^+] = c\alpha$ $= 0.15 \text{ x } 0.0205$ $= 3.\ 07$ $[H^+] = c\alpha$ $= 0.0015 \text{ x } 0.205$ $= 3.\ 08 \text{ x } 10^{-4}$
305	**Answers**	Chapter 10 Quick check 10.9 5 (a)  $[H^+] = c$ $= 0.12 \times 0.01041$ $= 1.25 \times 10^{-3}$	$[H^+] = c\alpha$ $= 0.12 \text{ x } 0.01041$ $= 1.25 \text{ x } 10^{-3}$
313	**Answers**	Chapter 10 STPM Practice 10 16 (a)  $PV = nRT$ $V = \frac{0.0474 \times 8.32 \times (35 + 273)}{1.01 \times 10^5} = 1.20 \times 10^{-3} \text{ m}^3$ $= 1.20 \times 10^{-3} \text{ m}^3$	$V = \frac{0.0474\ x\ 8.31\ x\ (35+273)}{1.01\ x\ 10^5}$
313	**Answers**	Chapter 10 STPM Practice 10 16 (b)(ii)  (ii) Titration of strong acid (HNO_3) with weak base (CH_3NH_2): Reaction: $CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3{}^+(aq) + OH^-(aq)$	CH_3NH_2 (aq) + HNO_3(aq) ⇌ $CH_3NH_3{}^+NO_3$(aq) OR CH_3NH_2 (aq) + H^+ (aq) ⇌ $CH_3NH_3{}^+$ (aq)
321	**Answers**	STPM Practice 12 Structured and Essay Questions 4. (b)  	$C_6H_{12}O_6$ (s) + $6O_2$ (g)
321	**Answers**	STPM Practice 12 Structured and Essay Questions 5 (a) (i)  Energy absorbed when 1 mol of CCl_4 bond break to form carbon and hydrogen gas atom under standard condition.	(a) (i) The enthalpy change when one mole of CCl_4 ia converted into its free gaseous atoms of Carbon and Chlorine under standard conditions.

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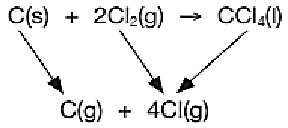
(ii)



$717.0 + 2(121.8) = -128.9 + x$

$x = +1090\ \text{kJ mol}^{-1}$

(iii) +800 kJ mol^{-1}, ΔH°_a of CCl_4 is higher th $SnCl_4$ as size of C is smaller than Sn bo strength C – Cl is stronger than Sn – Cl.

(b) $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$

Bond energy of $Cl_2 = 121 \times 2$

$= 242\ \text{kJ mol}^{-1}$

Mol of 1 molecule $= \frac{1}{NA} = 1.66 \times 10^{-24}$ mol

Energy to break 1 molecule of Cl_2

$= 1.66 \times 10^{-24} \times 242$ kJ

$= 4.02 \times 10^{-19}$ J

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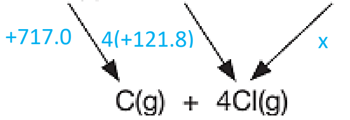
(ii)



$-128.9 + x = +717.0 + 4(+121.8)$

$x = +1533.1\ \text{kJ mol}^{-1}$

(iii) Size of C is smaller than Sn

Bond length of $SnCl_4$ is longer than of CCl_4 C – Cl is stronger than Sn-Cl

Enthalpy of atomisation of CCl_4 is higher than $SnCl_4$

(b)

$\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$

$\Delta H_{atm} = +121.0\ \text{kJ mol}^{-1}$

$Cl_2\ (g) \rightarrow 2Cl \quad 2 \text{ x } \Delta H_{atm}$

Bond energy = 2 (+121.0) = +242.0 kJ

Energy $= \frac{242.0}{6.02\ x\ 10^{23}} = 4.02 \text{ x } 10^{-22}$ kJ

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<td>322</td>
<td>Answers</td>
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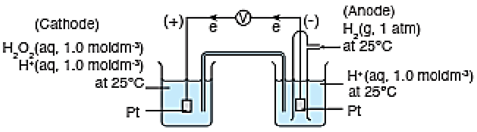
Chapter 13

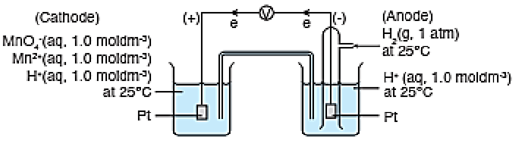
Quick check 13.2

2. (a) $MnO4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$

$E^o = +1.23$ V, more positive than SHE. Thus, half-cell of MnO^{4-}/Mn^{2+} act as cathode.



(b) $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l) \quad E^\circ = +1.77$ V more positive than SHE. Thus, half-cell of H_2O_2/H_2O act as cathode.



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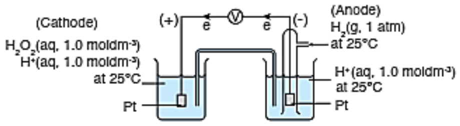
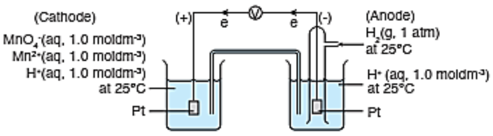
2. (b) $MnO4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$

$E^o = +1.23$ V, more positive than SHE. Thus, half-cell of MnO^{4-}/Mn^{2+} act as cathode.



(a) $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l) \quad E^\circ = +1.77$ V more positive than SHE. Thus, half-cell of H_2O_2/H_2O act as cathode.



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<td>324</td>
<td>Answers</td>
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Chapter 13

Quick check 13.4

5. $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

From Appendix, $E^\circ_{Sn^{4+}/Sn^{2+}} = +0.15$ V

$E = E^\circ - \frac{0.059}{6} \log \frac{[Sn^{2+}]}{[Sn^{4+}]}$

$+0.19 = +0.15 - \frac{0.059}{6} \log \frac{[Sn^{2+}]}{[Sn^{4+}]}$

$\frac{[Sn^{4+}]}{[Sn^{2+}]} = 22.6$

6. $Fe^{2+} + 2e^- \rightarrow Fe(s)$

From Appendix, $E^\circ_{Fe^{2+}/Fe} = 0.44$ V

$E = E^\circ - \frac{0.059}{6} \log \frac{1}{[Fe^{2+}]}$

$+0.46 = -0.44 - \frac{0.059}{6} \log \frac{1}{[Fe^{2+}]}$

$Fe^{2+} = 0.21\ \text{mol dm}^{-3}$

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5. $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

From Appendix, $E^\circ_{Sn^{4+}/Sn^{2+}} = +0.15$ V

$E = E^\circ - \frac{0.059}{2} \log \frac{[Sn^{2+}]}{[Sn^{4+}]}$

$+0.19 = +0.15 - \frac{0.059}{2} \log \frac{[Sn^{2+}]}{[Sn^{4+}]}$

$\frac{[Sn^{4+}]}{[Sn^{2+}]} = 22.6$

6. $Fe^{2+} + 2e^- \rightarrow Fe(s)$

From Appendix, $E^\circ_{Fe^{2+}/Fe} = 0.44$ V

$E = E^\circ - \frac{0.059}{2} \log \frac{1}{[Fe^{2+}]}$

$+0.46 = -0.44 - \frac{0.059}{2} \log \frac{1}{[Fe^{2+}]}$

$Fe^{2+} = 0.21\ \text{mol dm}^{-3}$

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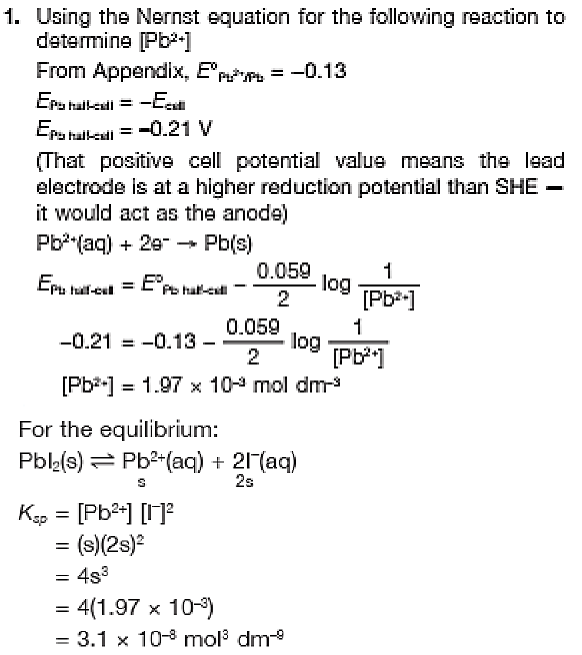
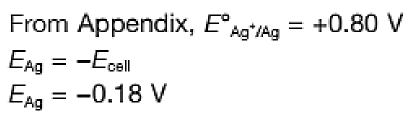
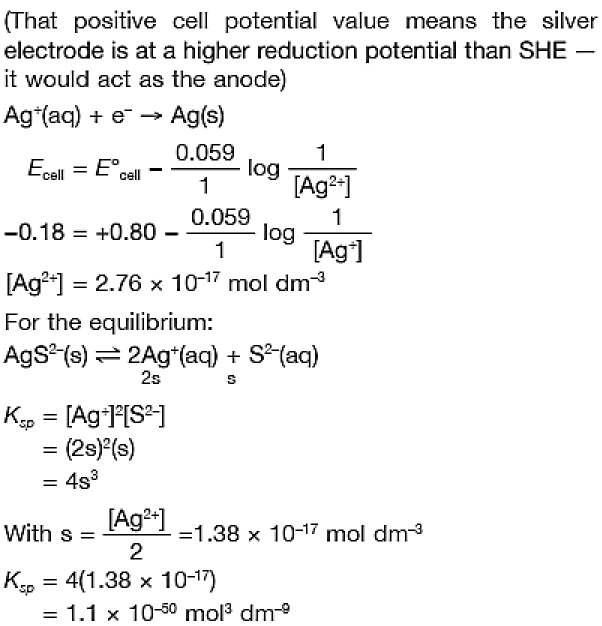
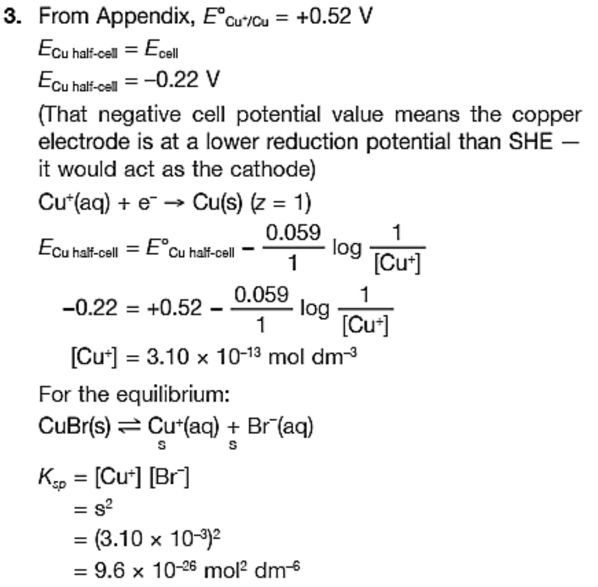
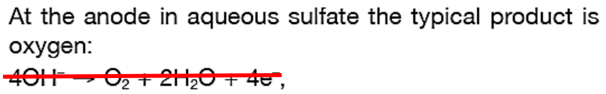
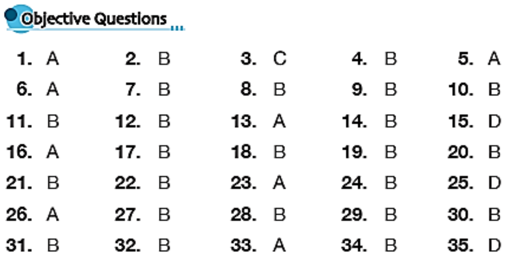
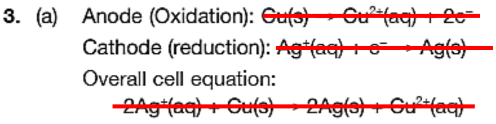
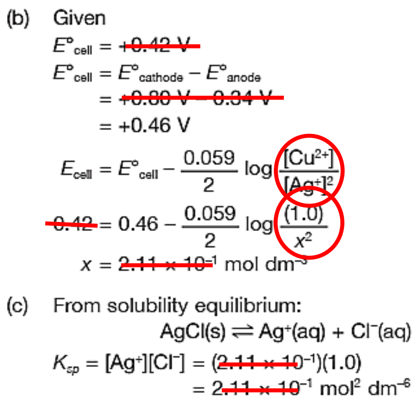
</table>

326	**Answers**	Quick Check 13.8  1. Using the Nernst equation for the following reaction to determine $[Pb^{2+}]$ From Appendix, $E^\circ_{Pb^{2+}/Pb} = -0.13$ $E_{Pb\ half\text{-}cell} = -E_{cell}$ $E_{Pb\ half\text{-}cell} = -0.21$ V (That positive cell potential value means the lead electrode is at a higher reduction potential than SHE – it would act as the anode) $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ $E_{Pb\ half\text{-}cell} = E^\circ_{Pb\ half\text{-}cell} - \frac{0.059}{2}\log\frac{1}{[Pb^{2+}]}$ $-0.21 = -0.13 - \frac{0.059}{2}\log\frac{1}{[Pb^{2+}]}$ $[Pb^{2+}] = 1.97 \times 10^{-3}$ mol dm^{-3} For the equilibrium: $PbI_2(s) \rightleftharpoons \underset{s}{Pb^{2+}(aq)} + \underset{2s}{2I^-(aq)}$ $K_{sp} = [Pb^{2+}][I^-]^2$ $= (s)(2s)^2$ $= 4s^3$ $= 4(1.97 \times 10^{-3})$ $= 3.1 \times 10^{-8}$ mol^3 dm^{-9}	Using the Nernst equation for the following reaction to determine $[Pb^{2+}]$ From Appendix, $E^o_{Pb^{2+}/Pb} = -0.13$ Since Pb has a more negative reduction potential, Pb acts as anode (oxidation): $Pb \rightarrow Pb^{2+} + 2e^-$ $E^\circ_{cell} = 0 - (-0.13) = 0.13$ V Using the Nernst equation: $E_{cell} = E^\circ_{cell} - \frac{0.059}{2}\log\frac{[Pb^{2+}]}{1}$ Substitute $E_{cell} = 0.21$: $0.21 = 0.13 - \frac{0.059}{2}\log[Pb^{2+}]$ $[Pb^{2+}] = 1.94 \times 10^{-3}\ moldm^{-3}$ For the equilibrium: $PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq)$ s 2s $K_{sp} = [Pb^{2+}][I^-]^2$ $= (s)(2s)^2$ $= 4s^3$ $= 4(1.94 \times 10^{-3})$ $= 2.9 \times 10^{-8}\ mol^3 dm^{-9}$
326	**Answers**	Quick Check 13.8  From Appendix, $E^\circ_{Ag^+/Ag} = +0.80$ V $E_{Ag} = -E_{cell}$ $E_{Ag} = -0.18$ V 2.  (That positive cell potential value means the silver electrode is at a higher reduction potential than SHE – it would act as the anode) $Ag^+(aq) + e^- \rightarrow Ag(s)$ $E_{cell} = E^\circ_{cell} - \frac{0.059}{1}\log\frac{1}{[Ag^{2+}]}$ $-0.18 = +0.80 - \frac{0.059}{1}\log\frac{1}{[Ag^+]}$ $[Ag^{2+}] = 2.76 \times 10^{-17}$ mol dm^{-3} For the equilibrium: $AgS^{2-}(s) \rightleftharpoons \underset{2s}{2Ag^+(aq)} + \underset{s}{S^{2-}(aq)}$ $K_{sp} = [Ag^+]^2[S^{2-}]$ $= (2s)^2(s)$ $= 4s^3$ With $s = \frac{[Ag^{2+}]}{2} = 1.38 \times 10^{-17}$ mol dm^{-3} $K_{sp} = 4(1.38 \times 10^{-17})$ $= 1.1 \times 10^{-50}$ mol^3 dm^{-9}	Since Ag has a more positive reduction potential, Ag acts as cathode (reduction): $Ag^+(aq) + e - \rightarrow Ag(s) \quad (z = 1)$ $E^\circ_{cell} = +0.80 - 0.00 = +0.80V$ $E_{cell} = E^o_{cell} - \frac{0.059}{1} log \frac{1}{[Ag^+]}$ $+0.18 = +0.80 - \frac{0.059}{1} log \frac{1}{[Ag^+]}$ $[Ag^+] = 3.10 \times 10^{-11}\ mol\ dm^{-3}$ For the equilibrium: $Ag_2S(s) \rightleftharpoons 2Ag^+(aq) + S^{2-}(aq)$ 2s s $K_{sp} = [Ag^+]^2[S^{2-}]$ $= (2s)^2(s)$ $= 4s^3$ With $s = \frac{[Ag^+]}{2} = 1.55 \times 10^{-11} mol\ dm^{-3}$ $K_{sp} = 4(1.55 \times 10^{-11})$ $= 1.5 \times 10^{-32}\ mol^3 dm^{-9}$

326	**Answers**	 **3.** From Appendix, $E^\circ_{Cu^+/Cu} = +0.52$ V $E_{Cu\ half\text{-}cell} = E_{cell}$ $E_{Cu\ half\text{-}cell} = -0.22$ V (That negative cell potential value means the copper electrode is at a lower reduction potential than SHE — it would act as the cathode) $Cu^+(aq) + e^- \rightarrow Cu(s)\ (z = 1)$ $E_{Cu\ half\text{-}cell} = E^\circ_{Cu\ half\text{-}cell} - \frac{0.059}{1} \log \frac{1}{[Cu^+]}$ $-0.22 = +0.52 - \frac{0.059}{1} \log \frac{1}{[Cu^+]}$ $[Cu^+] = 3.10 \times 10^{-13}$ mol dm^{-3} For the equilibrium: $CuBr(s) \rightleftharpoons \underset{s}{Cu^+(aq)} \underset{s}{+} Br^-(aq)$ $K_{sp} = [Cu^+]\,[Br^-]$ $= s^2$ $= (3.10 \times 10^{-3})^2$ $= 9.6 \times 10^{-26}$ mol^2 dm^{-6}	From Appendix, $E^o_{Cu^+/Cu} = +0.52\ V$ Since Cu has a more positive reduction potential, Cu acts as cathode (reduction): $Cu + (aq) + \ e - \ \rightarrow Cu(s)(z\ =\ 1)$ $E^\circ_{cell} = +0.52 - 0.00 = +0.52$ V $E_{Cu\ half-cell} = E^o_{Cu\ half-cell} - \frac{0.059}{1} log \frac{1}{[Cu^+]}$ $+0.28 = +0.52 - \frac{0.059}{1} log \frac{1}{[Cu^+]}$ $[Cu^+] = 8.55 \times 10^{-5}\ mol\ dm^{-3}$ For the equilibrium: $CuBr(s) \rightleftharpoons Cu^+(aq) + Br^-(aq)$ s s $K_{sp} = [Cu^+]\ [Br^-]$ $= s^2$ $= (8.55 \times 10^{-5})^2$ $= 7.3 \times 10^{-9}\ mol^2 dm^{-6}$
328	**Answers**	Chapter 13 Quick check 13.10 2.  At the anode in aqueous sulfate the typical product is oxygen: ~~$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$~~,	$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$
328	**Answers**	STPM Practice 13 Objective Questions  Objective Questions **1.** A **2.** B **3.** C **4.** B **5.** A **6.** A **7.** B **8.** B **9.** B **10.** B **11.** B **12.** B **13.** A **14.** B **15.** D **16.** A **17.** B **18.** B **19.** B **20.** B **21.** B **22.** B **23.** A **24.** B **25.** D **26.** A **27.** B **28.** B **29.** B **30.** B **31.** B **32.** B **33.** A **34.** B **35.** D	1. B 2. B 3. C 4. B 5. B 6. C 7. C 8. C 9. C 10. B 11. C 12. D 13. B 14. B 15. D 16. C 17. B 18. C 19. D 20. C 21. C 22. B 23. D 24. C 25. D 26. C 27. C 28. B 29. C 30. B 31. B 32. B 33. B 34. C 35. B
329	**Answers**	STPM Practice 13 Structured and Essay Questions 3(a)  **3.** (a) Anode (Oxidation): ~~$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$~~ Cathode (reduction): ~~$Ag^+(aq) + e^- \rightarrow Ag(s)$~~ Overall cell equation: ~~$2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$~~	(a) Anode: $Ag\ (s) \rightarrow Ag^+\ (aq) + e^-$ Cathode: $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$ Overall Equation: $2Ag\ (s) + Cu^{2+}\ (aq) \rightarrow 2Ag^+\ (aq) + Cu\ (s)$ (b) Given $E^\circ_{cell} = +0.26$ V $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ = +0.34 V – (+0.80 V) = - 0.46 V $E_{cell} = E^\circ_{cell} - \frac{0.059}{2} \log \frac{[Ag^+]^2}{Cu^{2+}}$ $0.26 = -\ 0.46 - \frac{0.059}{2} \log \frac{X^2}{(1.0)}$ $x = 6.89 \times 10^{-13}$ mol dm^{-3}

		 (b) Given E°_{cell} = ~~+0.42 V~~ $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ = ~~+0.80 V − 0.34 V~~ = +0.46 V $E_{cell} = E^\circ_{cell} - \frac{0.059}{2}\log\frac{[Cu^{2+}]}{[Ag^+]^2}$ ~~0.42~~ $= 0.46 - \frac{0.059}{2}\log\frac{(1.0)}{x^2}$ x = ~~2.11 × 10~~$^{-1}$ mol dm^{-3} (c) From solubility equilibrium: $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ $K_{sp} = [Ag^+][Cl^-]$ = (~~2.11 × 10~~$^{-1}$)(1.0) = 2~~.11 × 10~~$^{-1}$ mol^2 dm^{-6}	(c) $K_{sp} = [Ag^+][Cl^-] = (6.89 \times 10^{-13})(1.0)$ $= 6.89 \times 10^{-13}$ mol^2 dm^{-6}
330	**Answer**	6(a) The higher the , the greater the tendency for the redox reaction to proceed.	The higher the E°_{cell} , the greater the tendency for the redox reaction to proceed.