Task#76556

In a reaction it was found that 3.0g of a metal *X* was oxidized by 25.0 cm³ of 0.10 mol dm-³ K2 Cr2 O2 under acidic conditions.
(i) Deduce the mole of ratio between *X* and Cr2 O7²- ion.
(ii) write a balanced equation of the redox reaction.
(iii) Give the oxidation numbers of chromium and *X* in both their reduced and oxidized forms.[Molar mass of *X* = 200.6

Solution: Reduction:$Cr_2O_7^{2-} + 14\ H^+ + \mathbf{6}e = 2Cr^{3+} + 7H_2O$;

[Equivalent weight of $K_2\ Cr_2O_7$= molecular weight /**6** =M/6],

Where, M=Molecular weight of $K_2\ Cr_2O_7$

Oxidation: X $= X^{n+}$ +ne

[Equivalent weight of X = 200.6/n]

(i)Concentration of $K_2Cr_2O_7$ Solution =0.10 mol dm-³ = 0.1mol/lit = (0.1x M)g/lit =0.1x6 gm-eqivalent/lit= 0.6 Normal =0.6(N) ; [$1dm^3$ =1000cc = 1lit]

So, 1000cc 1(N) $K_2Cr_2O_7$ solution =200.6/n gm of X ;

1cc 1(N) $K_2Cr_2O_7$ solution =0.2006/n gm of X;

25cc 0.6(N) $K_2Cr_2O_7$ solution = 0.2006 x25x 0.6/n of X;

From given condition ,

$$0.2006 \text{ x}25\text{x } 0.6/n = 3$$

$$n = 0.2006\text{x}25\text{x}06/3 = 1.003$$

$$n = 1$$

Amount of $K_2Cr_2O_7$ in 25ml 0.10 mol/lit solution $=\frac{0.1\ X\ 25}{1000}$ mol = 2.5 $X10^{-3}$ mol

Amount of X in solution $=\frac{3g}{200.6g/mol}$ = 0.014955 mol

Mol ratio between X and $K_2Cr_2O_7$ $=\frac{0.014955}{2.5\ X10-3}$= 6;

That means for one mol of $K_2Cr_2O_7$ oxidises 6mols of metal X.

(ii) Reduction:$Cr_2O_7^{2-} + 14\ H^+ +6e = 2Cr^{3+} + 7H_2O$;

Oxidation: X $= X^+$ +e ;

Balanced equation of the redox reaction: $Cr_2O_7^{2-} + 14\ H^+ +6X = 2Cr^{3+} + 7H_2O + 6X^+$;

(iii) Oxidation number of Chromium in $Cr_2O_7^{2-}$ (oxidised form)=x=6, and in reduced form(Cr^{3+})=3

[2x +7(-2) =-2 ,or,x =6],metal (reduced form,X) =0, oxidised form (X^+)=1;