Clearly, the value of K_f (Formation Constant).

for the reaction

$$M^{2+} + L \underset{}{\overset{K_f}{\rightleftharpoons}} ML^{2+}$$ will be given by

$$K_f = \frac{[ML^{2+}]}{[M^{2+}][L]} \quad \text{---} (1)$$

In order to obtained the value of formation constant K_f, Solution containing known amounts of Total M^{2+} and Total 'L' are equilibrated. The absorption of these solutions at 550 nm is measured and the value of K_f is calculated as follows.

We know that

$$C_M = [M^{2+}] + [ML^{2+}] \quad \text{---} (2)$$

$$C_L = [L] + [ML^{2+}] \quad \text{---} (3)$$

and $$A = \epsilon_{(ML^{2+})} \cdot l \cdot [ML^{2+}]$$ {From Beer's Law}

$$\text{or } [ML^{2+}] = \frac{A}{\epsilon_{(ML^{2+})} \cdot l} \quad \text{---} (4)$$

Where, C_M = Total concentration of the metal ion
C_L = Total concentration of the ligand } Known

On putting the value of $[ML^{2+}]$ in eq^n (2) and (3) (which is obtained from eq^n 4) we get the values of $[M^{2+}]$ and $[L]$

Thus,
$$[M^{2+}] = C_M - \frac{A}{\epsilon_{[ML^{2+}]} \cdot l} \quad \text{---} (5)$$

$$[L] = C_L - \frac{A}{\epsilon_{[ML^{2+}]} \cdot l} \quad \text{---} (6)$$

In this way we get the value of $[ML^{2+}]$ $[M^{2+}]$ and $[L]$ and on putting these values in equation (i) we get the value of K_f