

# Comparative efficiencies of the degradation of C.I. Mordant Orange 1 using UV/H<sub>2</sub>O<sub>2</sub>, Fenton, and photo-Fenton processes

A.M. Gamal

Chemistry Department, Faculty of Science, Al-Azhar University (Girls), Nasr City, Cairo, Egypt.

**Abstract:** The kinetics of the photo degradation of C.I. Mordant Orange 1 imparted by UV/H<sub>2</sub>O<sub>2</sub>, Fenton and photo Fenton process were investigated. Negligible effects were observed in the presence of either UV light or H<sub>2</sub>O<sub>2</sub> alone. As expected the Fenton-mediated degradation occurred much faster than the photolytic process. Photo Fenton showed higher in treatment efficiency than that of Fenton process. The effects of dye concentrations, H<sub>2</sub>O<sub>2</sub> concentration, pH values and the presence of Fe<sup>2+</sup> concentration on the degradation rate constant were also studied. The rate constant of dye degradation that occurred in both the photolytic as well as in the Fenton and photo-Fenton processes was found to pseudo first-order kinetics.

[A.M. Gamal. Comparative efficiencies of the degradation of C.I. Mordant Orange 1 using UV/H<sub>2</sub>O<sub>2</sub>, Fenton, and photo-Fenton processes. Life Science Journal 2010;7(4):51-59]. (ISSN: 1097-8135).

**Keywords:** C.I. Mordant Orange 1, Degradation, Kinetic, UV/H<sub>2</sub>O<sub>2</sub>, Fenton, photo-Fenton processes.

## 1 Introduction

Hydrogen peroxide is increasingly favored as an environmentally acceptable bleaching agent both in domestic and industrial situations<sup>(1,2)</sup>. The combination of H<sub>2</sub>O<sub>2</sub> as a bleaching agent, and solar or UV radiation for photochemical degradation of textile dyes without using solid catalyst as titanium, iron, or manganese oxides, were the main objectives of many studies, H<sub>2</sub>O<sub>2</sub> is a safe, efficient and easy to use chemical oxidant suitable for wide usage on concentration prevention. Discovered by Thenard in 1818, it was first used to reduce odor in waste water treatment plants, and from then on, it became widely employed in waster water treatment<sup>(3)</sup>. However, since H<sub>2</sub>O<sub>2</sub> itself is not an excellent oxidant for many organic pollutants, it must be combined with UV light, salts of particular metals or ozone to produce the desired degradation results.

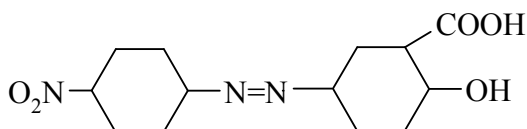
Many different types of synthetic dyes such as azo, vat, nitre, indigo, etc., are widely used for different purposes in paper and pulp manufacturing, plastics, printing and textile industry<sup>(4-5)</sup>. The effluent from these industries, when it contains substantial amounts of such dyes, causes not only coloration of water, but also poses a threat to aquatic life<sup>(6)</sup> furthermore; their presence in drinking water constitutes a potential human health hazard<sup>(7,8)</sup>. It is therefore essential either to remove the dyes from water or to treat them in such a way so as to minimize their effects on the environment and also to decolorize the water.

Different approaches have been suggested to remove dyes from aqueous solution including adsorption, biological degradation, coagulation, advanced oxidation processes (AOP), photo-Fenton reaction, ozone treatment and hypochlorite treatment<sup>(9-16)</sup>. A very simple approach which can be commonly utilized is- the photolytic oxidation of such solutions. In many cases, this method is readily applicable due to the UV content of sunlight. On the other hand, the degradation of dye using Fenton's reagent has been reported in the literature and has shown promising results<sup>(17-19)</sup>. In this paper, we wanted to examine the relative efficiencies of two different AOP methods, namely UV/H<sub>2</sub>O<sub>2</sub> and Fenton's for the degradation of the Alizaration yellow G dye, C.I. Mordant Orange 1 many of the parameters that affect dye degradation rate, such as dye concentration, H<sub>2</sub>O<sub>2</sub> concentration, and pH were also evaluated. Since many electrolytes are present in dye effluent, the kinetics of dye degradation in both the presence and absence of various ions were determined<sup>(20,21)</sup>.

## 2. Experimental

### 2.1. Reagents

C.I. Mordant Orange 1 (Fig. 1) was procured from CIBA-GEIGY (Switzerland) and was used as received distilled water was used for all dye solutions. Hydrogen peroxide (35%) was obtained from merck and ferrous sulphate heptahydrate (98%) was obtained from Merck.



C.I. Mordant Orange 1  
Mol. Wt. 287.00 g/mol  
Fig. (1)

## 2.2. Apparatuses

### 2.2.1. UV-visible spectroscopy

The absorption spectra are recorded with UV-9200 UV-Vis spectrophotometer. The absorbance of solutions measured at  $\lambda_{\max}$  (485 nm) with using a 1cm quartz cell.

### 2.2.2. Batch type photoreactor

All experiments are conducted in a 500 ml thermostated bath glass reactor equipped with a magnetic stirrer. The light source is low pressure mercury lamp (254 nm 2 lamp each of 4 watt).

### 2.2.3. pH measurements

The pH values of the solutions are adjusted using Multimeter; WTW (Wissenschaftlich-Technische werkstätten GmbH ) Imbolab Multi Lev 11, ba 12237 de Germany.

## 2.3. Preparation of solutions

### 2.3.1. Preparation of dye solutions

$5.0 \times 10^{-3}$ M stock solution of C.I. Mordant Orange 1 is prepared by dissolving 1.435gm of the dye in 1000 ml bi-distilled water from which working solutions are prepared ( $0.3-7.0 \times 10^{-5}$ M) by dilution.

### 2.3.2. Preparation of $H_2O_2$ solutions

Different concentrations of  $H_2O_2$  are freshly prepared (1.0, 2.0, 10.0, 30.0, 50.0, 70.0, mM). The concentrations of  $H_2O_2$  are large excess in comparison to dye solutions, so that  $H_2O_2$  concentration can be considered constant during the whole experiments.

### 2.3.3. Preparation of $Fe^{2+}$ solution

$1.0 \times 10^{-2}$  M freshly prepared stock solutions of  $FeSO_4 \cdot 7 H_2O$  are prepared and different concentrations ( $5.0 \times 10^{-6}$ , 1.0, 3.0, 5.0,  $7.0 \times 10^{-5}$ M) are prepared by dilution.

## 2.4. Homogenous photodegradation

The experiments are carried out in a bath-type photoreactor.  $H_2O_2$ , Fenton ( $Fe^{2+} + H_2O_2$ ) are acted as photocatalyst and UV light as illuminating light source. Reaction system is setup by adding the photocatalysts into 250ml dye solutions prepared in appropriate concentrations using bi-distilled water. The dye solutions are stirred and 5 ml samples are

withdrawn at regular time intervals and the dye concentrations are measured spectrophotometrically.

## 3. Results and discussion

### 3.1. Photocatalytic degradation with $H_2O_2$

#### 3.1.1. Effect of initial dye concentration

The effect of different concentrations of C.I. Mordant Orange 1 ( $0.3-7.0 \times 10^{-5}$ M) in presence of 30.0mM  $H_2O_2$  and pH 3.0 are studied. Initially, a large degree of removal is observed. This is due to fast decomposition of  $H_2O_2$  producing the hydroxyl radicals. Moreover, decolorization of dye is mainly due to hydroxyl radicals generated. Azo bonds are more active and C.I. Mordant Orange 1 contains one azo bond and decolorization of dye is due to the initial electrophilic cleavage of its chromophoric azo ( $-N=N-$ ) bond attached to naphthalene ring<sup>(22)</sup>.

The values of photodegradation Pseudo-first-order rate constant for different concentrations of C.I. Mordant Orange 1 calculated from the linear plots of  $\ln A/A_0$  against irradiation time<sup>(23)</sup> and the results of k are given in Table (1).

Taking into account that, the life-time of hydroxyl radical is very short (only few nanoseconds), they can only react where they are formed<sup>(24)</sup>. Increasing the quantity of C.I. Mordant Orange 1 molecules per volume unit logically enhance the probability of collision between organic matter and oxidizing species, leading to an increase in the degradation rate so that a rise in concentration induces an inner filter effect, i.e., incident light would largely be wasted for dye excitation rather than for the hydroxyl radical precursor excitation. Consequently, the solution becomes more and more impermeable to UV radiation. As the rate of hydrogen peroxide photolysis directly depends on the fraction of incident light absorbed by  $H_2O_2$  molecules, the degradation rate slows down.

#### 3.1.2. Effect of initial pH

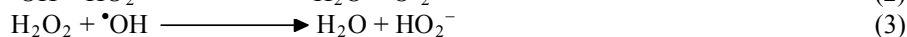
To study the effect of pH on photodegradation experiments are conducted for  $1.0 \times 10^{-5}$ M C.I. Mordant Orange 1 in presence of 30.0 mM  $H_2O_2$  at different initial pH values (2.03-8.15). The calculated pseudo first-order rate constants and the values of K are given in Table (2).

The results show that high degradation rate constant values are observed at pH 3.00. The alkaline medium, hydrogen peroxide undergoes decomposition leading to dioxygen and water rather than producing Hydroxyl radicals under UV irradiation<sup>(25)</sup>. Therefore the instantaneous concentration in  $^{\circ}OH$  is lower than that as expected.

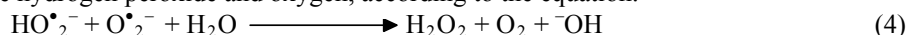
The base-catalyzed decomposition involves the  $\text{HO}_2^-$  anion: the conjugated base of  $\text{H}_2\text{O}_2$  reacts with a non dissociated molecule of  $\text{H}_2\text{O}_2$  according to the (eq. 1).



Furthermore, the deactivation of  $^-\text{OH}$  is greater when the pH of the solution is high (alkaline media), i.e. its expected that the reaction of  $^-\text{OH}$  with  $\text{H}_2\text{O}_2^-$  being approximately 100 time faster than its reaction with  $\text{H}_2\text{O}_2$ <sup>(26)</sup>.



The reactivity of  $\text{HO}_2^-$  and or its basic form  $\text{O}_2^{\bullet -}$  with organic compounds is very weak. They preferentially disproportionate producing some hydrogen peroxide and oxygen, according to the equation.



On the whole, the results demonstrate that it is very important to set a suitable pH in order to optimize the operating conditions.

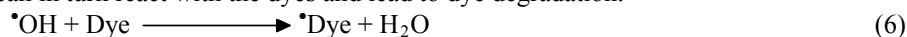
### 3.1.3. Effect of initial $\text{H}_2\text{O}_2$ dose on photo degradation .

The effect of varying the initial  $\text{H}_2\text{O}_2$  dose (1.0 – 70.0 mM) for C.I. Mordant Orange 1 concentration  $1.0 \times 10^{-5}$  and at pH 3.0 for the application of the UV/ $\text{H}_2\text{O}_2$  initiated degradation of dye effluent, it was necessary to initially change both the initial dye concentration and hydrogen peroxide concentration so as to optimize the conditions of maximum degradation.

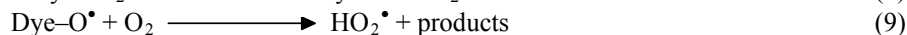
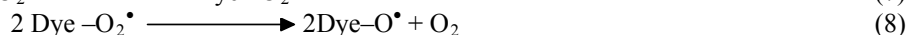
Experiments were carried out in the presence of either UV light or  $\text{H}_2\text{O}_2$  alone: neither of these treatments had an effect on the absorption maxima of the dye solution and imparted no degradation. However degradation of the dye occurred when it was subjected to UV light in the presence of  $\text{H}_2\text{O}_2$ , presumably due to the production of hydroxyl radicals, as shown below:



The hydroxyl radicals can in turn react with the dyes and lead to dye degradation.



Further reactions in solution can then take place as a part of the overall scheme:



Additionally the peroxy radicals can also further react with the dye as follows:

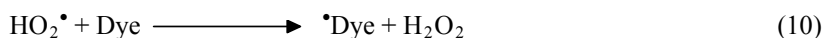


Table (3) shows the effect of increasing hydrogen peroxide concentrations on degradation rate with the most rapid degradation occurring at the highest  $\text{H}_2\text{O}_2$  concentration tested. However, it can be seen that the apparent rate seemed to be reaching at higher  $\text{H}_2\text{O}_2$  concentrations. For example, doubling the  $\text{H}_2\text{O}_2$  concentration from (1.0, 70.0 mM). This can be explained to be due to the self – to quenching reaction of  $^-\text{OH}$  radicals with  $\text{H}_2\text{O}_2$ .

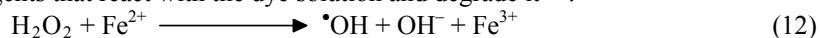


Therefore, 30.0 mM was selected as the optimum concentration of  $\text{H}_2\text{O}_2$  for subsequent study. Similarly, the initial concentration of the dye affected the degradation of the dye; Table (1) shows that as the dye concentration increased, an almost linear decrease in the degradation rate of the dye observed<sup>(27)</sup>. However, as we were unable to find an optimum concentration of C.I. Mordant Orange 1 that would give the best degradation of the dye, it therefore appeared that the apparent decoloriton rate was independent of dye concentration in the range of  $0.3-7.0 \times 10^{-5}$  mM.

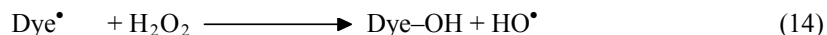
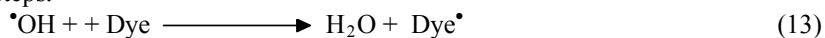
Examination of the effect of pH on dye degradation revealed that dye decoloration was affected on pH 3.0. In order to optimize the Fenton-mediated degradation of C.I. Mordant Orange 1 the concentrations of dye, hydrogen peroxide and  $\text{Fe}^{2+}$  were varied results are summarized Table (4-7) and Figs. (2a-5a).

### 3.2.: Fenton and photo-Fenton processes

The Fenton's reagent involves reaction of ferrous ions with hydrogen peroxide to produce hydroxyl radicals, which are strong oxidizing reagents that react with the dye solution and degrade it<sup>(28)</sup>.



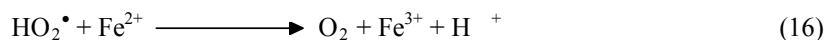
The hydroxyl radical propagates the reaction by reacting with the organic dye to produce further radicals, which can then react in many different steps.



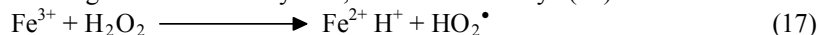
Additionally many other reaction are also possible, which include the radical-radical reaction or the reaction of the OH radical with  $\text{H}_2\text{O}_2$ .



The peroxide radicals ( $\text{HO}_2\cdot$ ) produced in the above case can further oxidize other species present in the solution<sup>(19)</sup>.



The rapid depletion of  $\text{H}_2\text{O}_2$  that is often observed with Fenton's reagents it probably due to the combined reactions (12-16). However, since reaction (13) has the highest rate constant, it is responsible for the degradation of the dye. Also, it is possible for  $\text{Fe}^{2+}$  to be auto regenerated in this system, and act as a catalyst(14).



Once can see from Table (4) that the rate C.I. Mordant Orange 1 degradation in the Fenton process was proportional to the amount of  $\text{H}_2\text{O}_2$ . This can be explained on the basis that at higher concentrations of  $\text{H}_2\text{O}_2$ , more OH radicals are available which can degrade more dye molecule. Similarly, at a fixed concentration of  $\text{H}_2\text{O}_2$ , the extent of dye degradation decreased with increasing concentration of the dye solution (Table 5). Lastly, Table (6) shows that as the amount of the  $\text{Fe}^{2+}$  increased, the apparent rate of dye degradation increased proportionally, presumably because of more OH radicals being generated.

It seems that unlike the photolytic process, the Fenton-mediated degradation of C.I. Mordant Orange 1 and  $\text{H}_2\text{O}_2$  concentrations. Lastly Table (7) shows that as pH: 3.0 the apparent rate of dye degradation is large but on increase the values of pH apparent rate is decrease. The photo-Fenton ( $\text{Fe}^{2+}/\text{UV}$ ) system is widely known Apos and is practical use<sup>(29)</sup>. The results are shown in Tables (4-7), and Figs. (2b-5b) illustrate that C.I. Mordant Orange 1 decolorization was excellent for photo-Fenton process instead of Fenton process. Due to the restriction of the power of UV-light used. The treatment efficiency improved slightly for dye waste water by photo-Fenton process. All pH values also dropped to acidic condition<sup>(30)</sup>.

### Conclusions

- The results obtained showed that  $\text{H}_2\text{O}_2/\text{UV}$  system could be efficiently used to degrade the C.I. Mordant Orange 1.
- Photodegradation efficiency of dye was negligible when photolysis was carried out in the absence of  $\text{H}_2\text{O}_2$  and UV light.
- The results indicated that the degree of degradation of C.I. Mordant Orange 1 was obviously affected by pH value and the amount. The optimum amount of Photocatalyst at pH: 3.0 and 30.0mM and at dye concentration of  $1.0 \times 10^{-5}\text{M}$ .
- It was observed that using of  $\text{Fe}^{2+}$  as a photocatalyst it increased the rate of degradation. The plot of  $\ln A/A_0$  dye versus irradiation time for C.I. Mordant Orange 1 was linear which showed that the photodegradation reaction obeys the pseudo-first-order kinetics reaction.
- The results show that the rate of reaction "k"  $\text{min}^{-1}$  in photo-Fenton large than Fenton.

**Table (1)** : Values of K ( $\text{min}^{-1}$ ) for the different concentration of the dye in presence of 30.0mM  $\text{H}_2\text{O}_2$  and pH 3.0 .

Concentration of dye M	K $\text{min}^{-1} \times 10^{-2}$	R
$0.30 \times 10^{-5}$	13.410	0.991
0.50	14.019	0.993
1.0	17.809	0.996
2.0	12.647	0.967
3.0	10.000	0.977
5.0	7.648	0.985
7.0	5.8954	0.995

**Table (2)** : Values of K ( $\text{min}^{-1}$ ) for pH in presence of  $1.0 \times 10^{-5}$ M C.I. Mordant orange 1, 30.0 mM  $\text{H}_2\text{O}_2$  .

pH	K $\text{min}^{-1} \times 10^{-2}$	R
2.03	5.243	0.957
3.00	8.112	0.932
4.00	6.210	0.961
5.10	4.250	0.952
7.09	3.198	0.960
8.15	2.609	0.981

**Table (3)** : Values of K ( $\text{min}^{-1}$ ) for the different concentration of the  $\text{H}_2\text{O}_2$  in presence of  $1.0 \times 10^{-5}$ M C.I. Mordant orange 1 and pH 3.0.

$\text{H}_2\text{O}_2$ mM	K $\text{min}^{-1} \times 10^{-2}$	R
1.00	0.0598	0.996
2.00	0.09184	0.996
10.00	0.2625	0.942
30.00	0.6432	0.956
50.00	0.4502	0.943
70.00	0.3200	0.956

**Table (4)** : Effect of  $\text{H}_2\text{O}_2$  concentration on C.I. Mordant orange 1 degradation in the Fenton, photo-Fenton processes .

Concentration of $\text{H}_2\text{O}_2$ mM	pH	Concentration of dye M	Concentration of $\text{Fe}^{2+}$ M	Fenton		Photo-Fenton	
				K $\text{min}^{-1} \times 10^{-2}$	R	K $\text{min}^{-1} \times 10^{-2}$	R
2.00	3.00	$1.00 \times 10^{-5}$	$3.00 \times 10^{-5}$	36.702	0.952	47.23	0.976
10.00				42.660	0.993	56.27	0.982
30.00				48.010	0.989	59.40	0.909
50.00				41.000	0.986	50.11	0.956
70.00				37.112	0.992	47.21	0.990

**Table (5)** : Effect C.I. Mordant orange 1 concentration on its degradation in the Fenton, Photo-Fenton Processes.

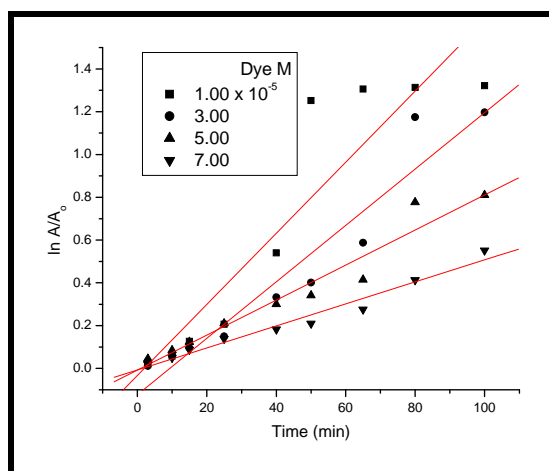
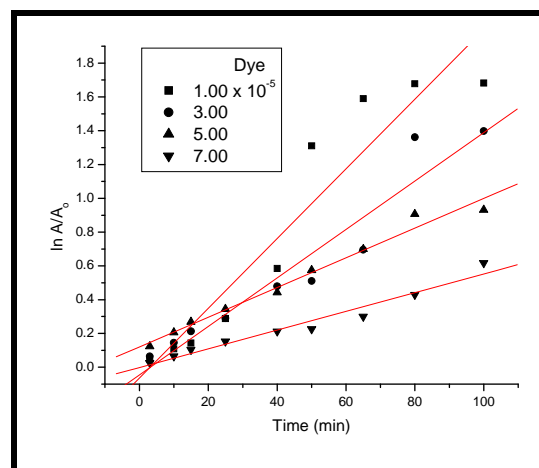
Concentration of dye M	Concentration of $\text{H}_2\text{O}_2$ mM	pH	Concentration of $\text{Fe}^{2+}$ M	Fenton		Photo-Fenton	
				K $\text{min}^{-1} \times 10^{-2}$	R	K $\text{min}^{-1} \times 10^{-2}$	R
$1.00 \times 10^{-5}$	30.00	3.00	$3.00 \times 10^{-5}$	16.67	0.952	35.72	0.986
3.00				15.84	0.954	28.16	0.994
5.00				13.66	0.964	20.90	0.996
7.00				5.51	0.990	8.44	0.943

**Table (6):** Effect of  $\text{Fe}^{2+}$  concentration on C.I. Mordant orange 1 degradation in the Fenton, photo-Fenton processes.

Concentration of $\text{Fe}^{2+}$ M	Concentration of $\text{H}_2\text{O}_2$	pH	Concentration of dye	Fenton		Photo-Fenton	
				$\text{K min}^{-1} \times 10^{-2}$	R	$\text{K min}^{-1} \times 10^{-2}$	R
$5.00 \times 10^{-6}$	30.00	3.00	$1.00 \times 10^{-5}$	1.034	0.959	2.66	0.996 0.972
$1.00 \times 10^{-5}$				4.239	0.972	7.14	
3.00				25.081	0.989	36.040	
5.00				13.721	0.953	20.66	
7.00				7.174	0.983	11.09	

**Table (7) :** Effect of pH on C.I. Mordant orange 1 degradation in the Fenton, photo-Fenton processes .

pH	Concentration of $\text{H}_2\text{O}_2$ mM	Concentration of dye M	Concentration of $\text{Fe}^{2+}$ M	Fenton		Photo-Fenton	
				$\text{K min}^{-1} \times 10^{-2}$	R	$\text{K min}^{-1} \times 10^{-2}$	R
3.00	30.00	$1.00 \times 10^{-5}$	$3.00 \times 10^{-5}$	14.227	0.988	19.723	0.980
4.00				9.261	0.981	11.166	0.969
5.00				5.130	0.973	8.589	0.996
6.00				2.810	0.984	5.634	0.973
7.00							

Fig. (2a) Kinetic of photodegradation of  $\text{Fe}^{2+}$   $3.0 \times 10^{-5}$  at different initial concentrations of C.I. Mordant orange 1 in presence of 30.00 mM  $\text{H}_2\text{O}_2$  in Fenton process and pH: 3.00.Fig. (2b) Kinetic of photodegradation of  $\text{Fe}^{2+}$   $3.0 \times 10^{-5}$  at different initial concentrations of C.I. Mordant orange 1 in presence of 30.00 mM  $\text{H}_2\text{O}_2$  in photo-Fenton process and pH: 3.00.

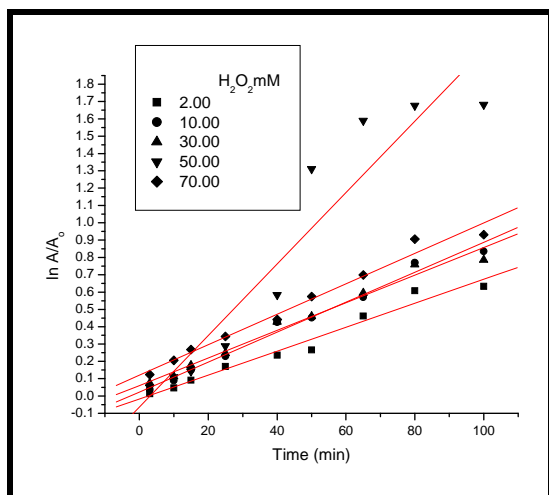


Fig. (3a): Kinetic of photodegradation of C.I. Mordant Orange 1  $1.0 \times 10^{-5}$  M at different initial concentrations of  $H_2O_2$  in presence of  $Fe^{2+}$ :  $3.0 \times 10^{-5}$  in Fenton process and pH: 3.00.

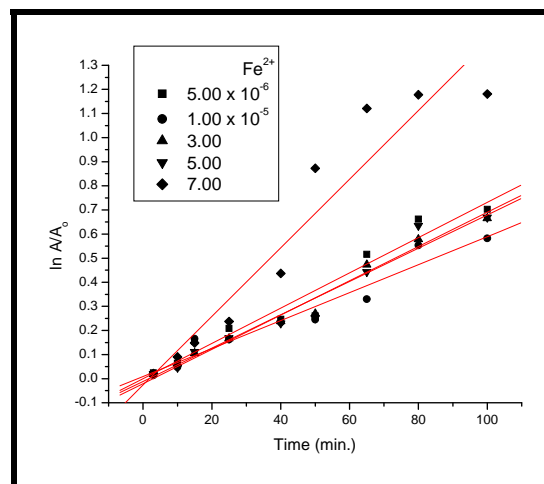


Fig. (4a): Kinetic of photodegradation of C.I. Mordant Orange 1  $1.0 \times 10^{-5}$  M at different initial concentrations of  $Fe^{2+}$  in presence of 30.00 mM  $H_2O_2$  Fenton process and pH: 3.00.

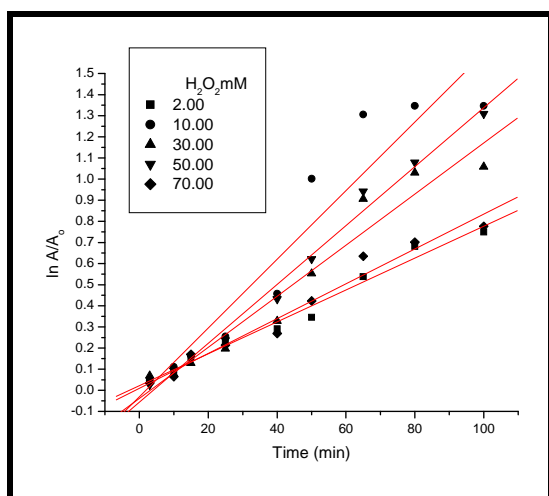


Fig. (3b): Kinetic of photodegradation of C.I. Mordant Orange 1  $1.0 \times 10^{-5}$  M at different initial concentrations of  $H_2O_2$  in presence of  $Fe^{2+}$ :  $3.0 \times 10^{-5}$  in Photo-Fenton process and pH: 3.00.

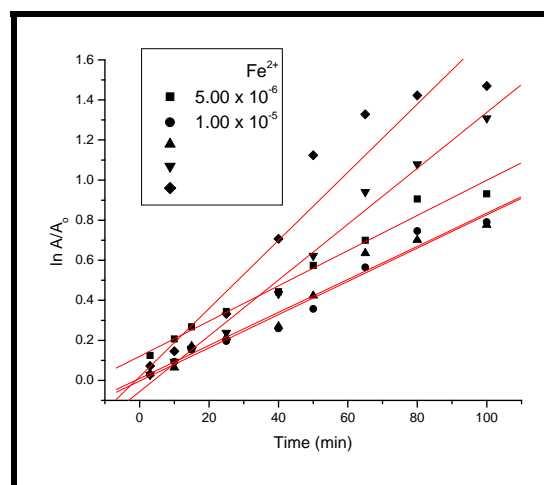


Fig. (4b): Kinetic of photodegradation of C.I. Mordant Orange 1  $1.0 \times 10^{-5}$  M at different initial concentrations of  $Fe^{2+}$  in presence of 30.00 mM  $H_2O_2$  photo-Fenton process and pH: 3.00.



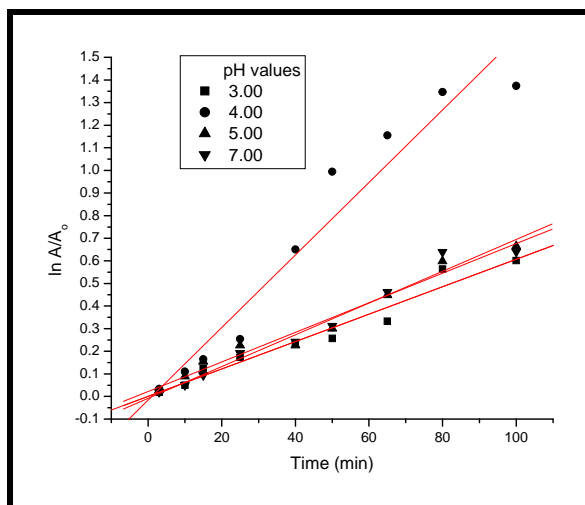


Fig. (5a): Kinetic of photodegradation of C.I. Mordant Orange 1  $1.0 \times 10^{-5}$  M at different initial concentrations of pH values in presence of 30.00 mM  $\text{H}_2\text{O}_2$  Fenton process and  $\text{Fe}^{2+}$   $3.0 \times 10^{-5}$  M

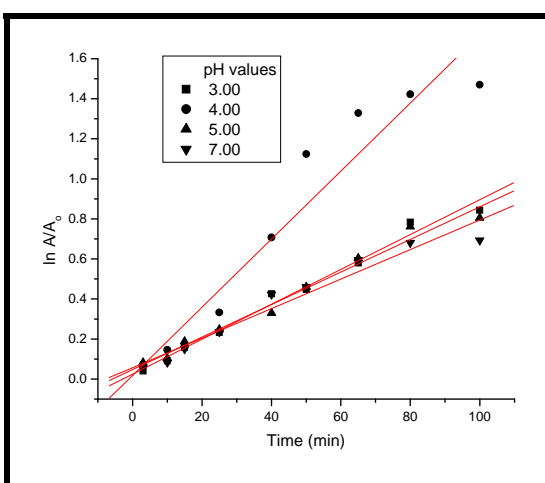


Fig. (5b): Kinetic of photodegradation of C.I. Mordant Orange 1  $1.0 \times 10^{-5}$  M at different initial concentrations of pH values in presence of 30.00 mM  $\text{H}_2\text{O}_2$  Photo-Fenton process and  $\text{Fe}^{2+}$   $3.0 \times 10^{-5}$  M

#### References

- Costa, F.A.P., Reis, E.M., Azeved, J.C. R., and Nozaki, J., *Solar Energy*, 77, 29-35 (2004).
- Gould, D.M., Griffith, W.P., Spiro, M., J. *Mol. Catal. A* 175, 289-291 (2002).
- Seheek, C.K., Frimmel, F.H., *Water Res.*, 29, 2346 (1995).
- Zollinger H. *Color chemistry*. New York: VCH Publishers; (1987).
- Neamtu M., Siminiceanu I. Yeldiler A. Kettrup A. *Dyes Pigments* (2002) 53:93.

- Nassar MM, Magdy YH. *Chem Eng J.* (1997) 66:223.
- Boeningo M. Carcinogenicity and metabolism of azodyes especially those derived from benzidine. DNHS (NIOSH) publication 80-119. Washington, D.C.: U.S. Gov. Printing Off.; July (1994).
- Pielesza A. *J. Mol Struct* (1999), 1-3:511.
- Sun Z., Chen Y., Ke Q, Yany Y, Yuan J. *Photochem Photobiol A*. 2002;149:169.
- Tanaka K, Padermpole K., Hisanga T. *Water Res.* (2000);34:327.
- Lin SH, Peng F.C. *Water Res.* (1996);30:587.
- Arslan I, Balcioglu Al. *Dyes Pigments* (1999) 43:95
- Perkowski J., Kos L. *Fibers Text East Eur* (2002);38:78.
- Solozhenko E.G., Soboleva N.M., Goncharuk VV. *Water Res.* (1995), 29:2206.
- Kang S.F., Lial CH. Chen M.C. *Chemosphere* (2002), 46(6):923.
- Rauf M.A., Ashraf S., Alhadrami S.N. *Dyes Pigments* (2005); 66:197.
- Swaminathan K., Sandhya S. Sophia A.C., K. Pachhade, Subrahmanyam Y.V. *Chemosphere* (2003),50:619.
- Tang W.Z., Chen. R.Z. *Chemosphere* (1996): 32: 947.
- Bae W. Lee SH, Ko G.B. *Water Sci Technol.* (2004) :49(4):91
- H.S. El-Desoky, M.M. Ghoneim, R. El-Sheikh, N.M. Zidan, Oxidation of Levafix GA reactive azo-dyes in industrial wastewater of textile dyeing by electrogenerated Fenton's reagent. *J. Hazard. Mater.* 175 (2010) 858-865.
- A. Fatema, Alshamsi, A.S., Albawda M.M., Alnuaimi, M.A. Rauf, S., Salman A., *Dyes and pigments* xx, 1-5 (2006).
- Murugandham, M. and Swaminatham, M., *Dye and Pigment*, 63. 315-32(2004).
- Sohrabi, M.R., Davallo, M.M., *Int. J. Chem. Tech. Res.* Vol. 1, No. 3, 446-451 (2009).
- Galindo, C., Jacques, P., and Kadt A., *J. Photochem. and Photobiol. A: Chem.* 141. 47 (2001).
- Hislop, K., and Bolton, J. *Environ. Sci. Technol.*, 33, 3119-3122 (1999).
- Rodriguez, M., Sarria, V., Esplugas, S., Pulgarin, C., J., *Photochem and Photobiol. A*, 151, 129-135 (2002)
- Patsoura, A., Kondarides, D.I. Very Kios, X.E., Enhancement of photoinduced Hydrogen production from Irradiated



- pt/TiO<sub>2</sub> suspensions with simultane degradation of Azo-dyes. Applied catalysis B: Environmental, 64, 171-179. (2007).
- 28- Ashraf S.S., Rauf M.A., Alhadrami S.N. Dyes Pigments (2006):69:80.
- 29- Esplugas, S., Rodriguez, M., Abderrazik, N., Conteras, S., Chammarro, E., Appl. Catal. B.: Environ., 37, 131-7 (2002).
- 30- Y.H. Huang, H.T. SU, L.W. Lin. Removal of citrate and hypophosfite binary components using Fenton, photo-Fenton and electro-Fenton processes. J. Environmental Sciences 21(2009) 35-40.

8/2/2010