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Adsorption of Phenol on Different Activated Carbons Prepared from Date Pits

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The potential use of different activated carbons (ACs) prepared from dates pits and phosphoric acid for the removal of phenol from aqueous solutions was investigated. Date pits were converted into five different types of activated carbons by air and phosphoric acid activation. The specific surface area (BET) of the prepared ACs varied from 794 m^2/g, for the phosphoric acid:date pit ratio of 5:1, to 1707 m^2/g for a ratio of 2:1. Batch adsorption experiments revealed that the adsorption of phenol varied among all of the prepared ACs, where the 2:1 AC showed the highest uptake. Equilibrium pH studies showed that the phenol removal was pH dependent and the maximum phenol uptake occurred at an equilibrium pH of 3.0. Dynamics studies indicated that the initial uptake of phenol on 2:1 AC at pH 4 was rapid, where 80% of the maximum uptake was achieved during the first 30 minutes; both surface adsorption and intraparticle diffusion were involved in the adsorption process and the data followed the pseudo second-order reaction. The equilibrium adsorption data of phenol on 2:1 AC at solution pH 3 was best described by the Redlich-Peterson, Sips, and Langmuir models.

Keywords activated carbon; adsorption; isotherm; phenol

INTRODUCTION

Phenol and substituted phenols are considered to be hazardous wastes, which are discharged into the aquatic environment from a vast number of sources, including oil refineries; synthesis of plastics and polymeric resins; leather and paper manufacture; coal conversion plants; petrochemicals; coal tar distillation; and pharmaceuticals, and paint production. These compounds are considered priority pollutants since they are harmful to organisms at low concentrations; many have been classified as hazardous pollutants because of their potential harm to human health. Chronic toxic effects in humans due to phenolic compounds include anorexia, vomiting, difficulty in swallowing, liver and kidney damage, headache, paralysis, and even death from respiratory failure or cardiac arrests (1). Fatal poisoning may also occur by adsorption of phenol through the skin, if exposure occurs over a large area of the body.

The typical concentration of phenols in waste water varies from 0.10 to 1600 mg/L (2). The decomposition of these compounds in nature is limited due to their stability and solubility in water; they have a half-life of between 2 and 72 days (3). It is, therefore, necessary to remove these compounds to levels below the maximum permissible concentration. The US Environmental Protection Agency (EPA) has set a limit of 0.1 mg/L of phenol in wastewater, while the World Health Organization (WHO) is stricter with regard to phenol regulation and sets a 0.001 mg/L limit on phenol concentration in potable water.

Several methods have been proposed to treat wastewaters containing phenolic compounds before they are discharged. These methods include biological degradation (4–6), chemical oxidation (7), ion exchange, solvent extraction (8–9), precipitation (10–11), and adsorption processes (12–16). For high concentration and low volume of phenolic wastewaters, adsorption of phenols on activated carbons (ACs) has been increasingly used to reach admissible standards (17–21). Activated carbons are versatile adsorbents that can be used in both liquid and gas phase adsorption processes due to their excellent physical and chemical properties. In general, these carbons have high adsorption capacities due to their highly developed pore structure characterized by large specific surface area. The porous structure consists of a network of interconnected macropores, mesopores, and micropores that provide a good capacity for the adsorption of organic molecules. The surface chemistry of these carbons, and the chemical

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characteristics of adsorbate, such as polarity, ionic nature, functional groups and solubility, determine the nature of the adsorption mechanism as well as the extent and strength of adsorption. A variety of physiochemical mechanisms/forces, such as Van der Waals, H-binding, dipole-dipole interactions, ion exchange, covalent bonding, and cation- and water-bridging, can be responsible for adsorption of organic compounds in activated carbon (22).

Agricultural by-products or wastes, such as saw dust, grain hulls, nut shells, date pits, peach pits, oat hulls, rice hulls, rice straw, coconut shells, almond shells, etc., provide an excellent source of activated carbon (17–21). Activated carbon from these agricultural sources can be prepared either by physical activation, which involves primary carbonization (below 700°C) followed by controlled gasification under the action of oxidizing gases at high temperature (up to 1100°C), or chemical activation where the precursor is mixed with a chemical that restricts the formation of tars (e.g., nitric acid, phosphoric acid, hydrogen peroxide, etc.). After kneading, the precursor is carbonized and washed to produce the final AC. The chemical incorporated into the interior of the precursor particles reacts with the thermal decomposition products reducing the evolution of volatiles and inhibiting the shrinkage of the particles. This increases the conversion of the precursor to carbon and produces carbons with large internal porosity. After the heat treatment, the chemical is removed by copious washing.

The present study investigates the use of different types of activated carbons prepared from date pits and phosphoric acid to remove phenol from aqueous solutions. The effect of solution pH on the adsorption isotherms was studied. The adsorption equilibrium data was analyzed using different two- and three-parameter isotherm models. Finally, the dynamics of adsorption of phenol on the prepared activated carbon was analyzed using the pseudo second-order kinetics and intraparticle diffusion.

EXPERIMENTAL

Materials and Adsorbents

The date pits used in the preparation of the activated carbons (ACs) were thoroughly washed with deionized water, dried to remove any dust or peels, then ground and sieved, where powder passing a standard 200-mesh screen was used in the activation.

The powdered date pits (C) were mixed with 85% phosphoric acid (A) in different weight ratios (A:C) of 1:2, 1:1, 2:1, 3:1, and 5:1 and diluted 5 times with deionized water in a 1000 mL beaker. The beaker was placed on the hot plate/stirrer and gently boiled for a few hours until the mixture turned to a black paste. The temperature of the paste at this stage was 160°C. In the same beaker, the heating continued while the material was stirred with a glass tube that introduced air at a flowrate of 3.0 L/min. This stage of carbonization and activation continued for 15 minutes and the temperature increased to approximately 215°C. The black paste (activated carbon) produced was poured into a test tube that was placed vertically in a tubular furnace. It was activated for 30 min with an airflow rate of 2.5 L/min at a temperature ranging from 300 to 500°C. After activation, the tube was removed from the furnace and cooled for 10 min with the flowing air.

The synthesized AC was washed by boiling in deionized water, filtered, and then rinsed with deionized water. This procedure was repeated twice, then the sample was placed in a soxhlet extractor and rinsed with deionized water for 48 hours. The rinsed carbon was transferred to a beaker and slurred with deionized water. The slurry was boiled and filtered. The wash water was tested with lead nitrate solution to ensure that all phosphate had been removed. After cooling, the pH of the wash water was measured to assure neutrality. The carbon was dried in a vacuum oven at 100°C for 12 hours, cooled in a dessicator, weighed, ground, transferred to a container, and stored in the dessicator (23).

Samples of the prepared ACs were tested using a Micrometrics ASAP 2010 gas adsorption surface area analyzer. In this test, nitrogen adsorption isotherms at 77 K were obtained and used to determine the surface area of the adsorbent using the BET equation (24). Micropore volume was determined using the *t*-plots; mesopore volume was determined using the BJH equation (25).

Determination of the Functional Groups on AC

The primary functional groups on the prepared ACs, such as carboxyls, lactones, lactols, and phenol groups, were found by the Boehm's titration method (23,26) as follows: 0.1 N solutions of sodium bicarbonate ($pK_a = 6.37$), sodium carbonate ($pK_a = 10.25$), and sodium hydroxide ($pK_a = 15.74$), were prepared. Two one-hundred mg portions of each of the prepared ACs were weighed and placed in a series of 250-ml Erlenmeyer flasks. Next, 50 mL of the above-mentioned bases were transferred to different flasks containing the carbon samples and empty flasks to be used as blanks for comparison. Samples and blanks were shaken at 150 rev./min for 48 hours at room temperature, and then left for 6 hours to settle the particulates. The clear solution was then filtered and 20 mL of filtrate was added to 15 mL of 0.1 N HCl. Excess HCl was determined by titration with 0.1 N NaOH. According to Boehm's titration method, sodium bicarbonate can neutralize carboxyl groups, sodium carbonate can neutralize carboxyl, lactone, and lactol groups, and sodium hydroxide can neutralize carboxyl, lactone, lactol, and phenol groups. Thus, the functional groups can be identified by the known volumes of acids and bases used in titration. Table 1 lists the different functional groups and relative concentration ($M_{eq}H^+/g$) of functional groups available on 2:1 AC.

TABLE 1
Concentration of surface functional groups on 2:1 AC

Group	($M_{eq}H^+/g$)
Lactones and Phenol	0.013
Lactones	0.005
Carboxyl group	0.005

Determination of Point of Zero Charge

The point of zero charge was determined using the acid–base titration method as described by Francisco et al. (29) Several 50-mL portions of 0.01 M NaCl solution were prepared in flasks; and the required pH for each portion was adjusted by the addition of the appropriate amount of 0.01 M solutions of either NaOH or HCl. When the required pH was achieved, 0.15 g of the prepared activated carbon sample was added to each flask and shaken for 24 hours, and the final pH was recorded. Control tests were conducted without the use of activated carbon samples to eliminate the influence of CO_2 on pH. The point of zero charge (pH_{pzc}) was defined as the point where the "pH final vs. pH initial" curve crosses the line "$pH_{initial} = pH_{final}$."

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Adsorption Experiments

Batch adsorption experiments were conducted by adding the appropriate amount of the prepared ACs into 100 mL reagent bottles containing 50 mL portions of phenol solution; differing initial concentrations (20–350 ppm) were used in the equilibrium studies. The mixtures in these bottles were agitated for a predetermined time in a shaker at 25°C. The phenol solution, then, was separated from the sorbent and the concentration of phenol was determined using an ultraviolet spectrophotometer (PU 8630 UV/VIS/NIR Single Beam) at a wavelength of 270 nm. The effect of varying pH was also studied by adjusting the pH in the range of 3.0 to 11.0 using diluted nitric acid or diluted sodium hydroxide. The adsorption capacity (i.e., uptake) was calculated using Eq. (1):

$$q_e = \frac{(C_0 - C_e)V}{w} \tag{1}$$

where q_e is the equilibrium uptake (mg/g), C_o is the initial phenol concentration (mg/L), C_e is the equilibrium phenol concentration (mg/L), V is the volume of the solution (L), and w is the mass of the adsorbent (g).

The dynamics studies were carried out by conducting batch adsorption experiments with 40 ppm of phenol at pH 4.0. Samples were taken after different time periods and analyzed for the phenol concentration.

RESULTS AND DISCUSSION

Characterization of Carbons

Table 2 shows the BET surface area, molecular volume, total pore volume, micropore volume, and pore diameter of the five prepared ACs. The SEM pictures of these ACs are shown in Fig. 1. Both the BET data and SEM images show that the surface area of the prepared ACs increases as the ratio of phosphoric acid to activated carbon increases, until it reaches the ratio of 2:1 (A:C), where it, then, starts to decrease. This trend can be attributed to the enlargement of the micropores until they are damaged by the effect of excessive amounts of the acid. Table 2 shows that acid activation enhanced the porosity of all ACs which resulted in a pronounced increase in the total pore volume and specific surface area for all activated carbons. These results can be explained on the basis of different activation mechanisms, namely the pyrolysis of the raw materials enhanced the cross linkage and, hence, created highly porous materials, while activation with phosphoric acid led to the elimination of water and, therefore, destruction of the cellulosic structure (28). The increase of the extent of activation from date pits to activated carbon was associated with the increase in the total pore volume and specific surface area and a decrease in the mean pore radius; suggesting the significant contribution of micropores to the total pore volume. However, in the case of activation for 2:1–5:1 ACs different results were observed, where the surface area and the total pore value was drastically decreased and the result for the pore diameter was ambiguous. This may be attributed to the fact that excess acid could form a protective solid layer that reduces the activation of the raw material. Due to the mixed micro/mesoporosity, the active carbons are not freely accessible to the probe molecules from aqueous solution. These findings are in an agreement with those by other researchers (28–29).

Adsorption of Methylene Blue on ACs

Adsorption of methylene blue on activated carbon is commonly used to test the adsorption properties, particularly the adsorption capacity. The resultant equilibrium isotherm for the adsorption of methylene blue on the 2:1 AC can be seen in Fig. 2. The results of adsorption of methylene blue on the various preparations of AC are shown in Fig. 3. The figure indicates that the prepared AC has a high adsorption capacity, even higher than many other AC preparations reported in the literature. Figure 3 also shows that the 2:1 AC exhibited greater adsorption capacity than the other AC preparations tested in the present study. This is likely due to the fact that the 2:1 AC preparation has a higher surface area than other AC preparations, as evident from Table 2.

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TABLE 2
Characterization of the prepared ACs

Sample ID	BET surface area (m^2/g)	VM (cm^3/g STP)	Single point adsorption		BJH adsorption average Pore diameter (4v/A)
			Total pore volume (cm^3/g)	Micropore volume (cm^3/g)	
1:2 AC	1078.56	247.76	0.520	0.4600	23.54
1:1 AC	1193.23	274.1	0.620	0.4800	20.87
2:1 AC	1706.69	392.05	1.140	0.5900	26.7
3:1 AC	1601.77	367.95	1.480	0.4700	36.91
5:1 AC	793.46	182.27	0.70	0.2400	35.44

Effect of pH on Adsorption of Phenol on 2:1 AC

Adsorption of aromatic compounds by activated carbons is a complex process. Many variables are involved in the adsorption process, such as solution properties and temperature; electrostatic interactions; dispersive and chemical interactions; and intrinsic properties of both the solute and adsorbent. Solution pH is one of the most important parameters affecting the adsorption of aromatic compounds on activated carbon. This is due to the fact that pH affects the solubility of phenol and the ionization state of the functional groups on the adsorbent.

Adsorption of phenol on 2:1 AC was investigated over the range of pH 3–11, and the results are shown in Fig. 4. The point of zero charge (pH_{PZC}) of the 2:1 AC was found to be 7.2 following the procedure outlined earlier in the experimental section. Figure 4 shows that the uptake of phenol on the AC reached a maximum at pH 3 and it decreased with the increase in solution pH. The decrease in phenol uptake was more pronounced at $pH > 7.5$. This is due to the dependency of phenol ionization on solution pH and the point of zero charge for mesoporous carbon, as well as the electrostatic and dispersive interactions between phenol and the activated carbon. While electrostatic interaction appears when the adsorbate is dissociated under the experimental conditions used, the dispersive interactions may be attributed to several possible mechanisms: π–π dispersion interactions, the hydrogen bonding formation mechanism proposed by Coughlin and Ezra (27), and the electron donor–acceptor complex proposed by Mattson et al. (28) The π–π dispersion interaction mechanism is the most widely accepted (29). Phenol is found in water solutions in the molecular state across a broad range of pH values. Increasing solution pH decreases the percentage of the unionized species and increases the percentage of ionized phenol. At low solution pH values, i.e., $pH < pH_{PZC}$, the carbon surface is positively charged and as there is no electrostatic repulsion between the unionized phenol and the positively charged surface; adsorption is higher. In this case, dispersive interactions are predominant, mainly because of the attraction between the π orbitals on the carbon basal planes and the electronic density in the phenol aromatic rings (π–π interactions). At pH values higher than pH_{PZC} the carbon surface is negatively charged, resulting in reduced adsorption due to the repulsive force between the phenolate ion and negative charge of the carbon. This occurs based on the relationship (27):

$$P_o = \frac{P_T}{(1 + 10^{pH - pK_a})} \tag{2}$$

where P_o is the concentration of unionized phenol species, P_T the total concentration of phenol taken, the pH, the

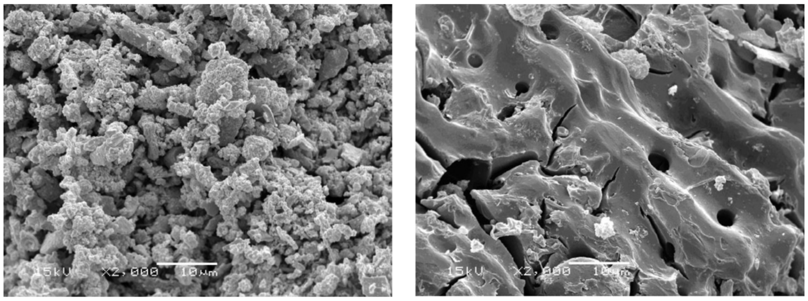

FIG. 1. SEM images of 1:2 AC (left) and 1:5 AC (right) preparations.

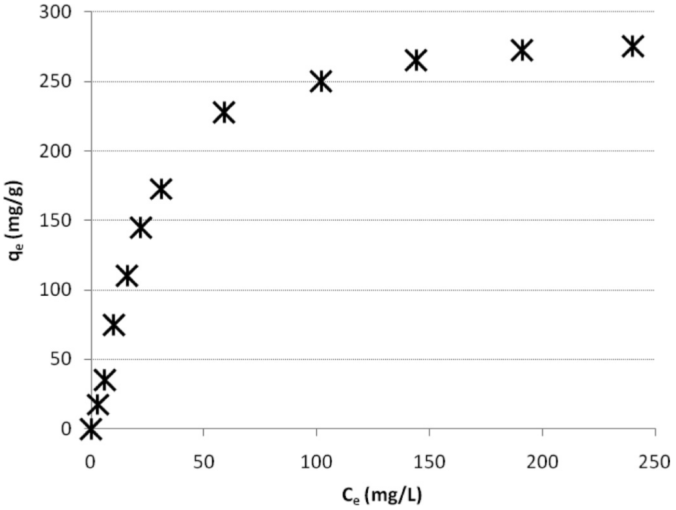

FIG. 2. Equilibrium isotherm for the adsorption of methylene blue on the 2:1 AC.

final pH (equilibrium pH) after adsorption, and pK_a is 9.95 for phenol. Thus, below pH 9.95, phenol will be found in solution predominantly in the molecular form; above pH 9.95, the ionic form predominates. Because the hydroxyl group is an activating group, the aromatic ring is partially negatively charged. It is worth mentioning that, despite the electrostatic repulsion at solution pH 11, a significant adsorption takes place, indicating a possibility for chemisorption involvement in the process. Similar results have been reported in other studies (21,36–38).

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Dynamics of Phenol Adsorption on 2:1 AC

The effect of contact time on phenol adsorption on 2:1 AC was investigated by determining the uptake of phenol from a 40 ppm solution at pH 4.0 at different time intervals (5 – 600 min); the results of the variation of phenol uptake versus time are presented in Fig. 5. The figure reveals that the uptake of phenol on 2:1 AC is rapid during the first 30 minutes, where more than 80% of the maximum uptake occurs during this period. After 30 minutes the adsorption process becomes slower, near the equilibrium. In between these two stages of uptake, the rate of adsorption is found to be nearly constant. This can be explained based on the fact that during the first 30 minutes, a large number of vacant surface sites are available for adsorption, thus the uptake increases rapidly. After 30 minutes, the remaining vacant surface sites are sterically and electrostatically hindered due to the repulsive forces between the solute molecules on the solid and bulk phases, thus the uptake increases more slowly. After 24 hours, the phenol uptake was found to increase by only 1 percent. Therefore, it was assumed that steady state kinetics was reached after 6 hours.

The adsorption of phenol on 2:1 AC was found to follow pseudo second-order kinetics. Analysis of the data was based on Eq. (3), as originally outlined by McKay (39),

$$\frac{dq_t}{dt} = k_2(q_e - q_t)^2 \tag{3}$$

where q_e (mg/g) is the uptake at equilibrium, q_t (mg/g) is the uptake at time t (min), k_2(g/mg · min) is the equilibrium rate constant of pseudo second-order sorption kinetics. Equation 3 can be solved with the boundary condition: $q = 0 \ @ \ t = 0$, to give Eq. 4.

$$\frac{t}{q_t} = \frac{t}{q_e} + \frac{1}{{q_e}^2 k_2} \tag{4}$$

The abovementioned parameters, i.e., q_e, and k_2 are determined by either fitting the equilibrium data to linear-equivalent forms or using nonlinear regression techniques.

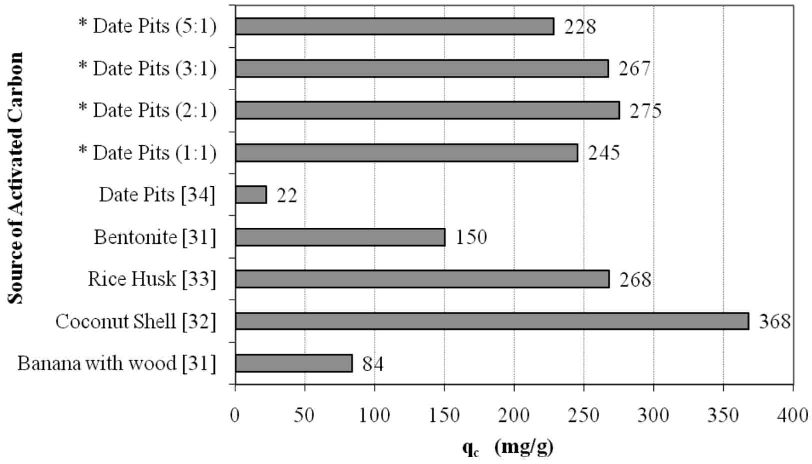

FIG. 3. Adsorption of methylene blue on different activated carbons, (*) represents data from the present study.

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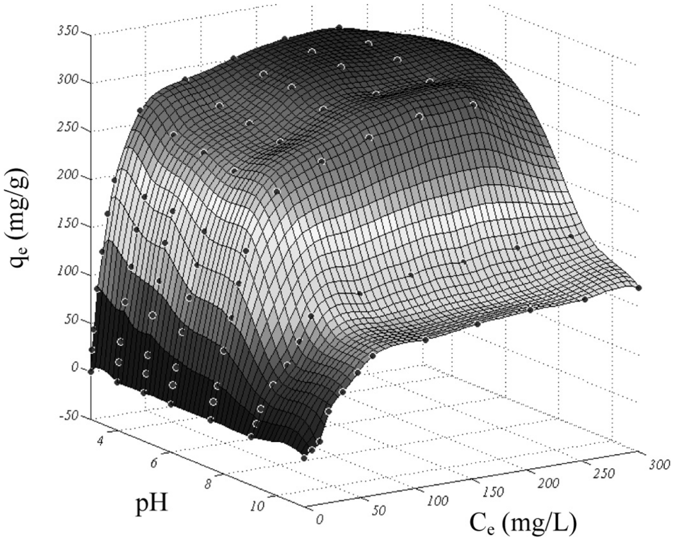

FIG. 4. Effect of pH on adsorption of phenol on the prepared 2:1 AC (mass of AC = 0.20 g).

In this study, the Composite Fractional Error Function (CFEF) (40), defined as seen in Eq. 5, was used.

$$\text{CPEF} = \min \sum_{i=1}^{p} \left[\frac{(q_{e.\exp} - q_{e.cal})^2}{q_{e.\exp}} \right]_i \tag{5}$$

Using CFEF, the parameters of the pseudo second-order rate Eq. (4), k_2 and q_e, for phenol solution of initial concentration of 40 ppm solution at pH 4.0, were found to be 0.001 g/mg · min and 83.4 mg/g, respectively. The low value of CFEF (1.4 mg/g) indicates that this model appropriately describes the kinetics of adsorption for the prepared ACs.

Weber and Morris (41) proposed a very simple equation, (6), below, that can be used to investigate the contribution of intraparticle diffusion in the overall adsorption process,

$$q_t = k_{id} t^{0.5} + C \tag{6}$$

where, k_{id} (mg/g · min$^{1/2}$) is the intraparticle diffusion rate constant and C (mg/g) is a constant that is related to the thickness of the boundary layer, i.e., the larger the value of C the greater is the boundary layer effect (21,37). According to this equation, if particle diffusion is involved in the sorption process, then a plot of uptake versus the square root of time reveals a linear relationship. Particle diffusion would be the rate controlling step if this line passes through the origin. However, if the data exhibit multi-linear plots, then two or more steps influence the sorption process. As elucidated in Fig. 6 for the adsorption of phenol on 2:1 AC, the results show that "multilineratity" can be assumed, suggesting that the overall adsorption process is controlled by many steps, e.g., film or external diffusion, pore diffusion, and surface diffusion and adsorption on the pore surface. The first step is characterized by a steep change in the uptake, which may indicate an instantaneous adsorption stage resulting from external surface adsorption. The second stage is characterized by a moderate increase in the uptake, suggesting intraparticle / pore diffusion within the pores of the AC. The third stage exhibits slow adsorption, indicating that equilibrium has been reached, where the intraparticle diffusion starts to slow down due to extremely low soluble concentration in the solution AC.

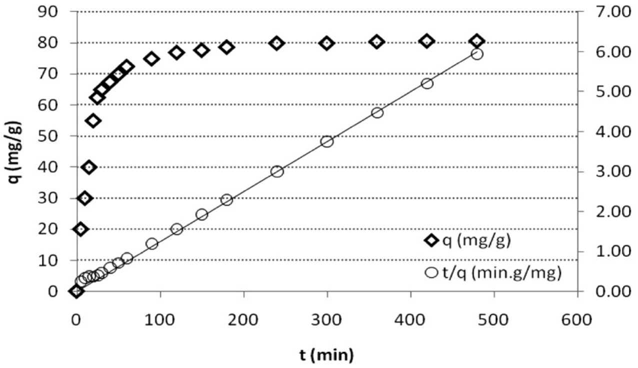

FIG. 5. Effect of contact time on adsorption of phenol on the prepared 2:1 AC (pH = 4.0, mass of AC = 0.20 g).

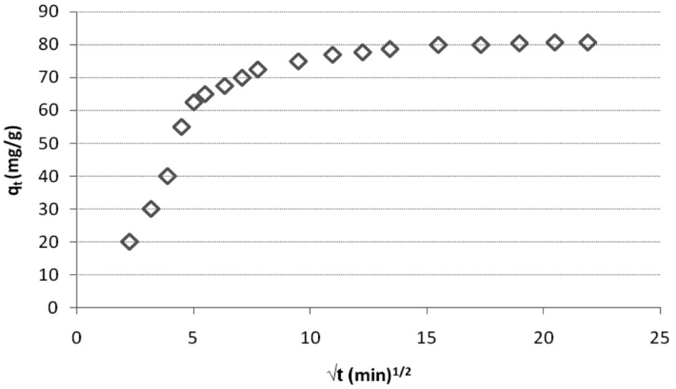

FIG. 6. Application of Weber and Morris equation on adsorption of phenol on the prepared 2:1 AC (pH = 4.0, mass AC = 0.20 g).

ADSORPTION EQUILIBRIUM

The effect of equilibrium concentration of phenol on the adsorption capacity of the different prepared ACs is presented in Fig. 7. Of these, 2:1 AC, at pH 3 showed the maximum adsorption capacity, with an adsorption of 315.5 mg/g. The equilibrium phenol concentration provides an important driving force to overcome all mass

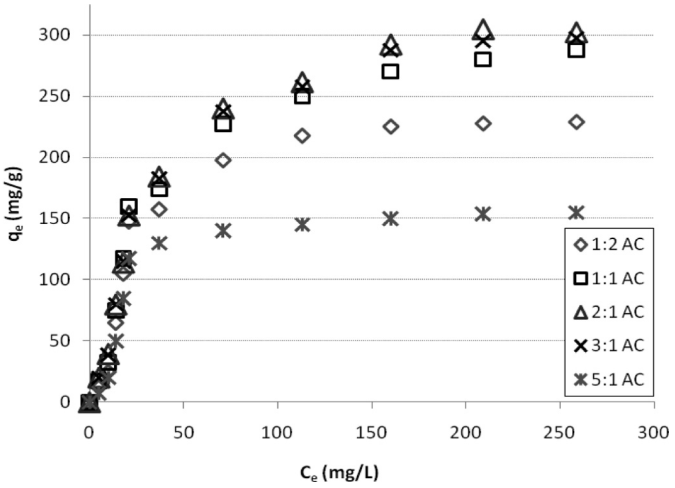

FIG. 7. Equilibrium isotherm of adsorption of phenol on the prepared ACs (pH = 3.0, mass of AC = 0.20 g).

transfer resistances of phenol between the aqueous and solid phases, thus a higher equilibrium concentration of phenol on the AC. The high sorption capacity of 2:1 AC was due to the high surface area. Figure 8 lists the adsorption capacity of different adsorbents for phenol. As can be seen in this figure, the adsorption capacities of our prepared ACs are among the highest.

For the analysis of adsorption data, models that describe the adsorption isotherm data are needed. Four such models, namely the Langmuir, Freundlich, Sips, and Redlich-Peterson isotherm models, were used to analyze the adsorption of phenol on 2:1 AC at solution pH 3.

The Langmuir isotherm model (7):

$$q_e = \frac{q_{mon} K_L C_e}{1 + K_L C_e} \tag{7}$$

is considered one of the most widely used models to describe and analyze adsorption equilibrium isotherms. The adjustable parameters q_{mon} and K_L are the Langmuir constants, which are related to the amount of adsorption corresponding to monolayer coverage, or adsorption capacity, and the energy of adsorption, respectively. This model is valid for monolayer sorption onto a surface with a finite number of identical sites and uniform adsorption energies where there are interactions between the adsorbed molecules.

The Freundlich isotherm model (8):

$$q_e = K C_e^{1/n} \tag{8}$$

is another widely used equation to describe adsorption data. This empirical model is based on the assumption that sorption occurs on heterogeneous surfaces. However, it does not provide any information on the monolayer adsorption capacity. The adjustable parameters K and n are the Freundlich constants and are indicators of adsorption capacity and adsorption intensity, respectively.

The Sips model, (9), is an example of a three-paramenter adsorption isotherm equation:

$$q_e = \frac{q_{mon} K_L (C_e)^{1/n}}{1 + K_L (C_e)^{1/n}} \tag{9}$$

where q_{mon}, K_L, and n are the Sips constants. At low sorbate concentration, this model assumes the form of the Freundlich model, while at high concentrations it predicts a constant, monolayer sorption behavior similar to that of the Langmuir isotherm.

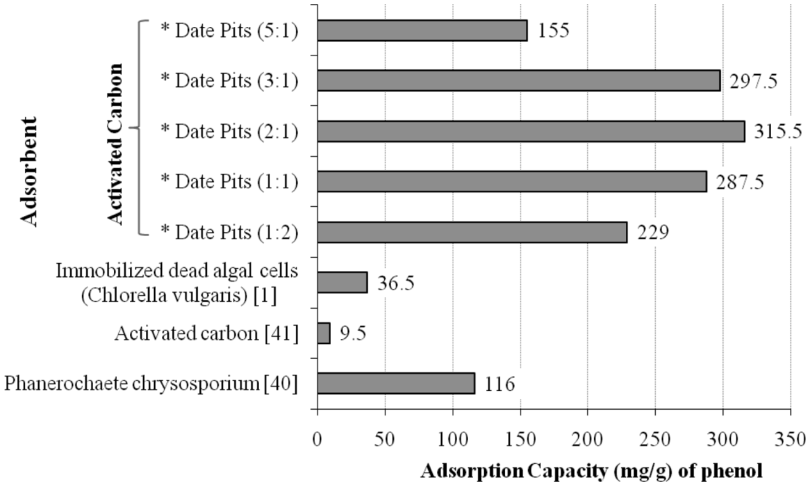

FIG. 8. Sorption capacities for phenol using different adsorbents, (*) represents data from the present study.

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TABLE 3
Isotherm parameters for the sorption of phenol on the prepared 2:1 AC

Model	Parameter	Value
Langmuir	q_{mon} (mg/g)	344.2
	K_L (L/mg)	0.0558
	CFEF (mg/g)	4.3
Freundlich	K $(l/mg)^{1/n}$(mg/g)	35.0
	n	2.26
	CFEF (mg/g)	24.6
Sips	q_{mon} (mg/g)	335.4
	K_L $(L/mg)^{1/n}$	0.048
	n	0.92
	CFEF (mg/g)	3.1
R-P	K_1	17.75
	K_2	0.046
	B	1.05
	CFEF (mg/g)	2.5

Another three-parameter model that is used to analyze adsorption data is the Redlich-Peterson (R-P) isotherm model, which is given by Eq. 10,

$$q = \frac{K_1 C_e}{1 + K_2 C_e^B} \tag{10}$$

where K_1, K_2, and B are the R-P constants. It should be noted that for values of B equal to one, the R-P equation reduces to the Langmuir equation.

The Composite Fractional Error Function (CFEF), defined by equation 5, was used as a non-linear error function to determine the parameters of these models. The parameters, along with the values of CFEF, are listed in Table 3. The values of CFEF indicate that the R-P, Sips, and Langmuir isotherm models could fit the adsorption of phenol on the prepared ACs well, with a little advantage of R-P model, as illustrated graphically in Figure 9.

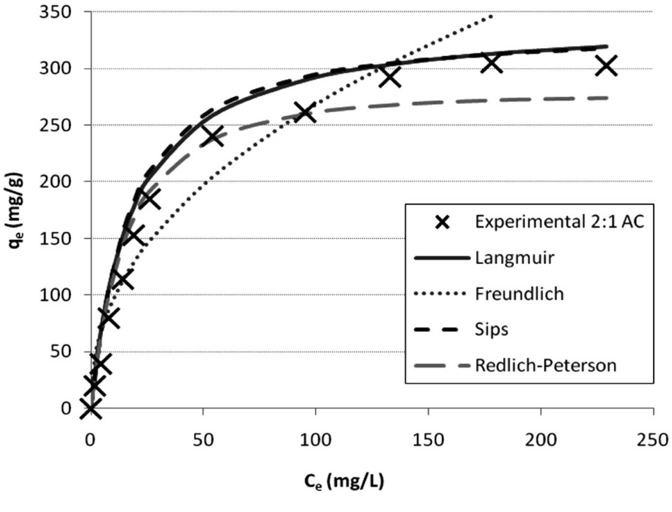

FIG. 9. Equilibrium isotherm of adsorption of phenol on 2:1 AC.

CONCLUSIONS

Date Pits have proven to be able to produce activated carbons with high surface area and, thus, high adsorption capacity. Adsorption of phenol on the prepared activated carbons was found to be pH dependent; it decreased by increasing pH. Dynamics studies of phenol on the prepared activated carbons revealed that it followed pseudo second-order kinetics with the involvement of film, pore and surface diffusions, as well as adsorption on the pore surface. Equilibrium studies indicated that Redlich-Peterson, Sips, and Langmuir isotherm models best described the adsorption of phenol on the prepared activated carbon.

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