a) At this point, [HA] concentration is given and no OH⁻ has been added yet. So this is the starting point of the titration. Use the $K_a$ reaction because this is a problem of HA dissociating slightly in water to make some $H^+$, and we want to know pH.

$$\begin{align*}
\text{HA (aq)} & \quad + \quad \text{H}_2\text{O (l)} & \rightleftharpoons & \quad \text{A}^- (aq) & \quad + & \quad \text{H}_3\text{O}^+ (aq) \\
\text{Initial} & \quad 0.200 \text{ M} & & 0 & & 0 \\
\text{Change} & \quad -x & & +x & & +x \\
\text{Equilibrium} & \quad 0.200 - x \approx 0.200 & & x & & x
\end{align*}$$

Using the equilibrium constant expression:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \approx \frac{(x)(x)}{(0.200)} = \frac{x^2}{0.200} = 1.9 \times 10^{-4} \quad \text{(which was given as the value for } K_a)$$

Solving for $x$, we get $x = \sqrt{(0.200)(1.9 \times 10^{-4})} = 6.16 \times 10^{-3} = [\text{H}_3\text{O}^+]$

Therefore, $\text{pH} = - \log [\text{H}_3\text{O}^+] = 2.21$

Note that the small $x$ approximation is justified because $x$ turned out to be on the order of $10^{-3}$, so subtracting $x$ from 0.200 would not make a significant change in the 0.200.

b) You can recognize this as the titration midpoint because 0.200 moles of HA were initially present and now 0.100 moles of OH⁻ have been added, so it’s halfway to adding 0.200 moles of OH⁻ which would be the titration equivalence point.

At the titration midpoint, $\text{pH} = pK_a$ (note: if this were a base being titrated instead, and if you knew the $K_b$ value, the fastest way to calculate pH would be from $\text{pOH} = pK_b$ at titration midpoint)

$$\text{pH} = pK_a = - \log(1.9 \times 10^{-4}) = 3.72$$

c) You can recognize this as the equivalence point because exactly the same amount of OH⁻ has been added to equal the amount of HA that was initially present. $[\text{HA}] = 0.200 \text{ M}$ and $[\text{OH}^-] = 0.200 \text{ M}$. Since OH⁻ is strong, it reacts entirely, taking away the same amount of HA, to make A⁻. This describes the initial solution. Then A⁻ hydrolyzes slightly to make some OH⁻ at equilibrium.
Use the $K_b$ reaction because you have HA, $A^-$ and OH$^-$ as the chemicals about which you know some information about concentrations. The solution described by the initial conditions above (with 0.200 M of $A^-$ in it) is exactly as if you created a solution by adding 0.200 moles of $A^-$ to 1 L of water.

\[
\begin{array}{c|c|c|c}
\text{A}^- (aq) & + & \text{H}_2\text{O} (l) & \rightleftharpoons \\
0 & & 0.200 \text{ M} & 0.200 \text{ M} \\
\text{Initial} & +0.200 \text{ M} & - 0.200 \text{ M} & - 0.200 \text{ M} \\
& = 0.200 \text{ M} & = 0 & = 0 \\
\text{Change} & - x & + x & + x \\
\text{Equilibrium} & 0.200 - x \approx 0.200 & x & x \\
\end{array}
\]

This reaction has an equilibrium constant of $K_b$, which you can get from knowing $K_a$ by using the equation that $K_aK_b = K_w$.

\[
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.9 \times 10^{-4}} = 5.26 \times 10^{-11} \text{ (keep one more sig fig than needed for now, until we round later)}
\]

\[
K_a = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} \approx \frac{(x)(x)}{(0.200)} = \frac{x^2}{0.200} = 5.26 \times 10^{-11}
\]

Solving for $x$, we get $x = \sqrt{(0.200)(5.26 \times 10^{-11})} = 3.24 \times 10^{-6}$, which is the OH$^-$ concentration.

So, $p\text{OH} = – \log [\text{OH}^-] = 5.49$, and $p\text{H} = 14 – p\text{OH} = 8.51$. Note that this number has 2 sig figs (two digits after the decimal place) because the smallest number of sig figs in the given information was 2 sig figs in the value for $K_a$ that was given.

Note that the small $x$ approximation is justified because $x$ turned out to be on the order of $10^{-6}$, so subtracting $x$ from 0.200 would not make a significant change in the 0.200.

d) You can recognize this as after the equivalence point because more OH$^-$ has been added than the amount of HA initially present. The excess OH$^-$ present is the amount beyond the equivalence point. Initially, there were 0.200 moles of HA present. So if 0.300 moles of OH$^-$ have been added, there is an excess of 0.100 moles of OH$^-$. There is no need to set up the ICE table because $x$ will be small, and adding $x$ to the 0.100 moles of OH$^-$ will not change the 0.100 moles of OH$^-$, so there is no need to calculate $x$.

\[
[\text{OH}^-] = \frac{0.100 \text{ moles OH}^-}{1.00 \text{ L}} = 0.100 \text{ M}
\]

\[
p\text{OH} = – \log [\text{OH}^-] = 1.000
\]

\[
p\text{H} = 14 – p\text{OH} = 13.000
\]
e) The titration curve must be a graph of pH on the y-axis and amount of titrant (NaOH) added on the x-axis. Parts a, b, c and d above give four points (red diamonds) on this S-shaped curve, which are indicated in the graph below. The rest of the graph (blue line) is a sketch that connects these points to make an S-shaped curve which has an inflection point at the equivalence point of the titration.