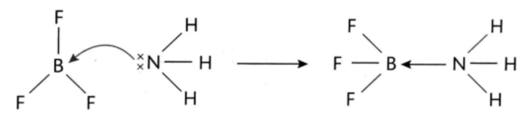
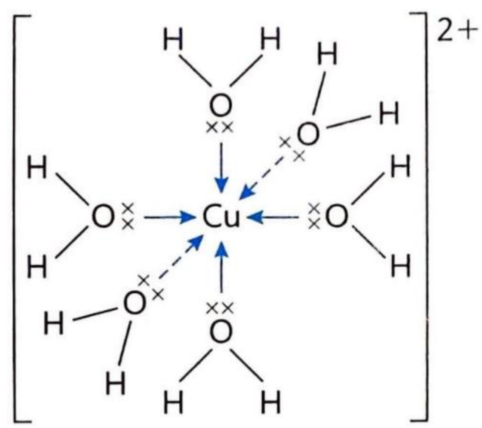
Name: ______________________ Date: __________ Period: ______

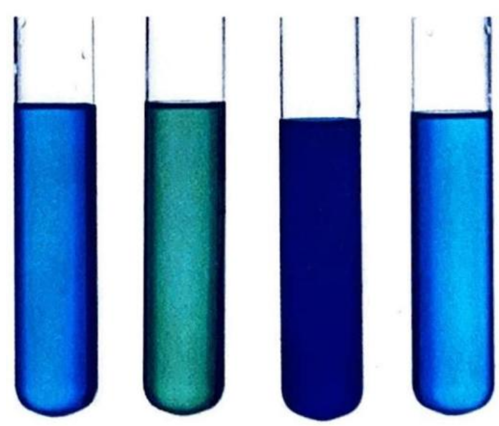
HLTopic 8_ Acids_Bases - Guided Notes

Slide 3: Lewis Acids and Bases

- Molecules with _____________ octets are often Lewis acids
- **Coordinate bond**
- _____________ metals are often Lewis acids (ions have incomplete d orbitals) that bond with ligands to form complex ions
- **$Cu^{2+}(aq)$ + $6H_2O(l) \rightarrow [Cu(H_2O)_6]^{2+}(aq)$**
- _____________ complex ions of copper form distinct colors

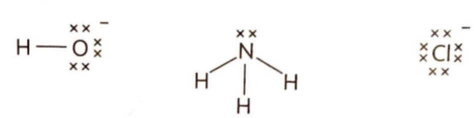


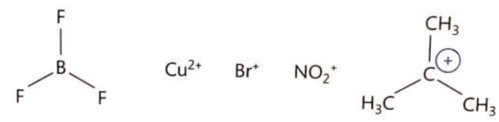




Slide 4: Nucleophiles and Electrophiles

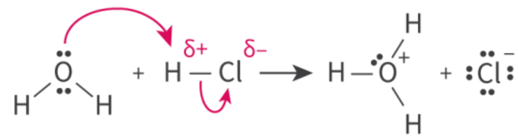
- **_____________:** electron-rich species that donates a lone pair to form a new covalent bond in a reaction
- **_____________:** electron-deficient species that accept a lone pair from another reactant to form a new covalent bond
- These are Lewis acids
- These are Lewis bases





Slide 5: Lewis Acids and Bases

- All _____________-Lowry acids are Lewis acids, but not all Lewis acids are Bronsted-Lowry acids
- Lewis base
- Lewis acid
- Lewis acid
- Lewis base
- Al_2O_3 (s) + NaOH(aq) → $NaAlO_2$(aq) + $2H_2O$(l)
- +
- Lewis acid
- Lewis base
- Some Lewis acids/bases are amphoteric but not _____________



Al_2O_3 (s) + 6HCl (aq) ⟶ $2AlCl_3$ (aq) $3H_2O$ (l)

Slide 6: You Try!

19 For each of the following reactions identify the Lewis acid and the Lewis base.

(a) $4NH_3(aq) + Zn^{2+}(aq) \rightarrow [Zn(NH_3)_4]^{2+}(aq)$
(b) $2Cl^-(aq) + BeCl_2(aq) \rightarrow [BeCl_4]^{2-}(aq)$
(c) $Mg^{2+}(aq) + 6H_2O(l) \rightarrow [Mg(H_2O)_6]^{2+}(aq)$

20 Which of the following could not act as a ligand in a complex ion with a transition metal?

A Cl^- **B** NCl_3 **C** PCl_3 **D** CH_4

21 Which of the following reactions represents an acid–base reaction according to Lewis theory but not according to Brønsted–Lowry theory?

A $NH_3(aq) + HCl(aq) \rightleftharpoons NH_4Cl(aq)$
B $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
C $Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons [Cu(NH_3)_4]^{2+}(aq)$
D $BaO(s) + H_2O(l) \rightleftharpoons Ba^{2+}(aq) + 2OH^-(aq)$

Slide 7: Dissociation Constants

- **_____________ constant that describes the extent to which a substance ionizes in solution**
- Used to describe _____________ of weak acids and bases in solution

- Since weak acids and bases do not dissociate completely, the ____________ of ions in solution **cannot be deduced from initial concentrations**
- Value of the dissociation constant tells information about the position of equilibrium (higher means more dissociation/higher ____________ of products compared to reactants).

Slide 8: Dissociation Constants

- **Acid ____________ constant (K_a):** quantitative measure of extent of dissociation of a weak acid in solution
- $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$
- $K_a = [H_3O^+][A^-]$
- [HA]
- **Base ____________ constant (K_b):** quantitative measure of extent of dissociation of a weak base in solution
- $B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$
- $K_a = [BH^+][OH^-]$
- [B]
- K_a and K_b are constant at a specified ____________
- Higher K_a indicates a stronger acid (____________ more)
- Higher K_b indicates a stronger base (____________ more)
- Acids that are ____________ (release more than one hydrogen ion) have multiple K_a values

Slide 9: Let's Practice!

- Write the expressions for K_a and K_b for the following acid and base:
- $CH_3COOH(aq)$
- $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
- $NH_3(aq)$
- $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
- $K_a = [CH_3COO^-][H^+]$
- $[CH_3COOH]$
- $K_b = [NH_4^+][OH^-]$
- $[NH_3]$

Slide 10: You Try!

- Write the expressions for K_a and K_b for the following:
- $C_2H_5NH_2(aq)$
- $HSO_4^-(aq)$

- CO_3^{2-}(aq)
- 2. Place the following acids in order of increasing strength:
-   H_3PO_4 $K_a = 7.1 \times 10^{-3}$
-   HNO_2 $K_a = 7.1 \times 10^{-4}$
-   H_2SO_3 $K_a = 1.2 \times 10^{-2}$
- 3. Why do you think we do not usually use the concept of K_a and K_b to describe the strength of strong acids and bases?

Slide 11: Calculations Involving Ka and Kb

- **Things to know:**
- K_a and K_b can be used to calculate ion ____________ at equilibrium
- pH/pOH values refer to H^+/OH^- ____________ **at equilibrium**
- The given ____________ of an acid or base is the **initial** concentration (before dissociation)
- ____________ used in Ka and Kb values must be **equilibrium concentrations**
- When K_a and K_b values are small, it can be assumed that $[acid]_{initial} \approx [acid]$________ $[base]_{initial} \approx [base]_{equilibrium}$
- ICE tables are used to calculate K_a/K_b and/or equilibrium ____________

Slide 12: Calculation of Ka and Kb from pH and initial concentration

- Calculate K_a at 298 K for a 0.01 mol dm^{-3} solution of ethanoic acid (CH_3COOH). It has a pH of 3.4 at this ____________.
- $[H^+]$ at ____________: $10^{-3.4} = 4.0 \times 10^{-4}$ mol dm^{-3}
-   $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
- ____________ 0.01 0.00 0.00
- Change -4.0×10^{-4} $+ 4.0 \times 10^{-4}$ $+ 4.0 \times 10^{-4}$
- ____________ $0.01-4.0 \times 10^{-4}$ 4.0×10^{-4} 4.0×10^{-4}
- $K_a = [CH_3COO^-][H^+]$
-   $[CH_3COOH]$
- $= (4.0 \times 10^{-4})^2$
-   0.01
- ~0.01
- $= 1.6 \times 10^{-5}$

Slide 13: Calculation of [H+]/pH, [OH-]/pOH from Ka and Kb

- A 0.75 mol dm^{-3} solution of ethanoic acid has a value for $K_a = 1.8 \times 10^{-5}$ at a specified temperature. What is its pH at this temperature?

- $[H^+]$ at equilibrium: $10^{-3.4} = 4.0 \times 10^{-4}$ mol dm^{-3}
- $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
- Initial 0.75 0.00 0.00
- Change -x + x + x
- Equilibrium 0.75 -x x x
- $K_a = [CH_3COO^-][H^+]$
- $[CH_3COOH]$
- $= x^2$
- 0.75
- ~0.75
- $= 1.8 \times 10^{-5}$
- $x = [H^+] = \mathbf{3.7 \times 10^{-3}}$
- pH $= -\log(3.7 \times 10^{-3}) =$ **2.4**

Slide 14: You Try!

- The acid dissociation constant of a weak acid has a value of 1.0×10^{-5} mol dm^{-3}. What is the pH of a 0.1 mol dm^{-3} aqueous solution of HA?
- Calculate the K_b of ethylamine, $C_2H_5NH_2$, given that a 0.10 mol dm^{-3} solution has a pH of 11.86.
- What are the $[H^+]$ and $[OH^-]$ in a 0.10 mol dm^{-3} solution of an acid that has $K_a = 1.0 \times 10^{-7}$?

Slide 15: pKa and pKb

- Similar to the ____________ between $[H^+]/[OH^-]$ and pH/pOH, K_a and K_b can be converted into their negative logarithms, known as pK_a and pK_b
- $pK_a = -\log(K_a)$ $K_a = 10^{-pKa}$
- $pK_b = -\log(K_b)$ $K_b = 10^{-pKb}$

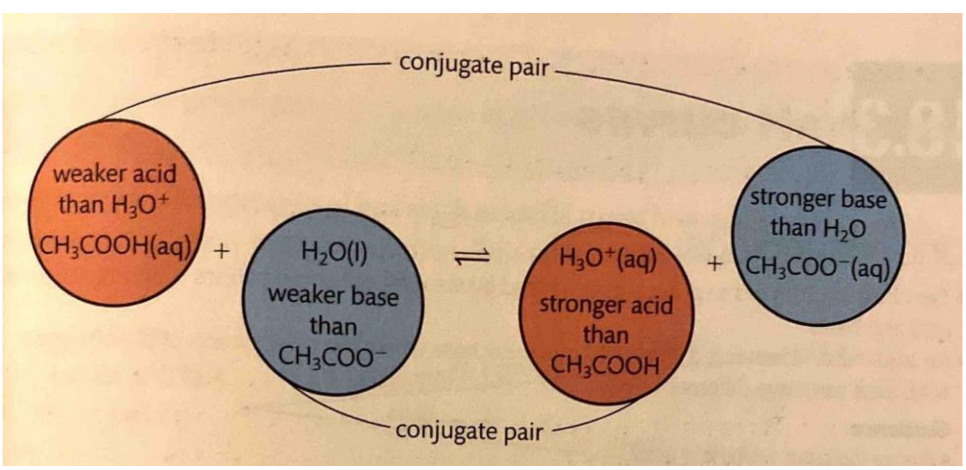
Slide 16: Important Things to Remember

- 1. pK_a and pK_b numbers are usually ____________ and have no units.
- 2. The ____________ between K_a and pK_a and between K_b and pK_b is inverse.
- 3. A change in one pK_a/pK_b unit ____________ a 10-fold change in the K_a/K_b
- 4. pK_a and pK_b must be quoted at a specific ____________

Slide 17: Relationship between Ka/Kb and pKa/pKb for a conjugate pair

- $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$
- $A^-(aq) + H_2O(l) \rightleftharpoons HA(aq) + OH^-(aq)$
- $K_a = [H^+][A^-]$
- [HA]
- $K_b = [HA][OH^-]$
- $[A^-]$
- $K_a \times K_b = [H^+][A^-]$
- [HA]
- $[HA][OH^-]$
- $[A^-]$
- ×
- $K_a \times K_b = [H^+]\times[OH^-] = K_w = 1.00 \times 10^{-14}$
- $pK_a + pK_b = pK_w = 14.00$ at 298 K
- Higher values of K_a mean lower values of K_b for its ____________ base
- Stronger acids have weaker conjugate bases (____________ lies more to the right)

Slide 18: Relationship between Ka/Kb and pKa/pKb for a conjugate pair



Slide 19: You Try!

- What is the relationship between K_a and pK_a?
- The pK_a of HCN is 9.21 and that of HF is 3.17. Which is the stronger acid?
- What are the pK_b values of CN^- and F^- from problem 2? Which is the stronger base?
- (a) The pK_a of ethanoic acid, CH_3COOH, at 298K is 4.76. What is the pK_b of its conjugate base. CH_3COO^-?
- (b) The pK_a of methanoic acid, HCOOH, at 298K is 3.75. Is its conjugate base weaker or stronger than that of ethanoic acid?

Slide 20: Buffers

- Acts to reduce the impact of one thing on another
- Water is vulnerable to fluctuations in pH
- Has major impacts on chemical reactions in aqueous solutions
- Biochemical systems are dependent on buffers due to enzymes needing to operate within a narrow pH range
- Examples: Mammalian blood, oceans
- Chemical processes also use buffers
- Examples: Electrophoresis, fermentation, dyes industry, calibration of instruments
- Different buffers can be made to buffer at almost any pH
- **A solution resistant to changes in pH when small amounts of acid or alkali are added**

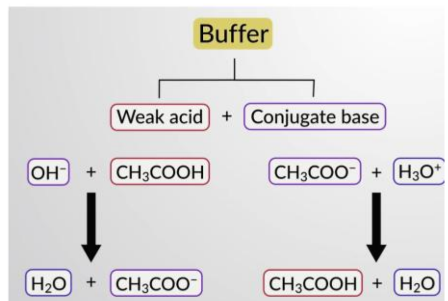
Slide 21: The ocean acts as a CO2 sink, absorbing CO2 from the atmosphere As more CO2 is absorbed, the water becomes more acidic/releases more H+ ions This causes equilibrium to shift and reduce CO32- levels

- **Ocean Water Buffer System**
- **Ocean Acidification:**
- ____________ carbonate levels have caused a decline in coral reefs
- Decreasing water pH causes a decreased buffering capacity and reduces ocean ability to absorb ____________ CO_2
- $CO_3^{2-} + H^+ \rightleftharpoons HCO_3^-$

Slide 22: How Buffers Work

- **Acidic buffers** ____________ pH < 7
- **Basic buffers** ____________ pH > 7
- Formed by mixing a weak acid/base with a solution of its ____________ acid/base

Slide 23: Acidic Buffers



- Example: Buffer made with $CH_3COOH(aq)$ and $NaCH_3COO(aq)$
- $NaCH_3COO(aq) \rightarrow CH_3COO^- + Na^+(aq)$
- $CH_3COOH(aq) \rightleftharpoons CH_3COO^- + H^+(aq)$
- Equilibrium lies to the left due to CH_3COOH being a weak acid
- This soluble salt dissociates fully
- Mixture contains high concentrations of both CH_3COOH and CH_3COO^-
- The acid and its conjugate base react with added OH^- and H^+ in neutralization reactions

Slide 24: Basic Buffers

- Example: Buffer made with $NH_3(aq)$ and $NH_4Cl(aq)$
- $NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$
- $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
- Equilibrium lies to the left due to NH_3 being a weak base
- This soluble salt dissociates fully
- Mixture contains high concentrations of both NH_3 and NH_4^+
- The base and its conjugate acid react with added OH^- and H^+ in neutralization reactions

Slide 25: How to Make a Buffer

- Start with an acid or base that has a pK_a or pK_b value close to ____________ pH/pOH of the buffer
- Mix this with a solution of a salt ____________ its conjugate or partially neutralize the solution with a strong acid/base
- The ____________ reaction ensures about half of the starting acid or base is converted into salt
- Final mixture contains equal amounts of the weak acid/base and the salt of its ____________ acid/base
- pH of buffer ____________ on:
- pKa or pKb of its acid/base
- The ratio of the initial ____________ of acid/salt or base/salt

Slide 26: Dilution

Ka and Kb are not changed by dilution

pH is not changed by dilution as the ratio of acid/base to salt does not change

Dilution affects amount of acid/base it can absorb without changes in pH (buffering capacity).

Temperature

Does affect Ka/Kb and pH

- **Factors that ____________ buffers**

Slide 27: Let's Practice!

- State whether each of the following mixtures will form a buffer solution when dissolved in 1.00 dm^3 of water:
- 0.20 mol $NaHCO_3$ and 0.20 mol Na_2CO_3
- 0.20 mol CH_3COOH and 0.10 mol HCl
- 0.20 mol NH_3 and 0.10 mol HCl
- 0.10 mol H_3PO_4 and 0.20 mol NaOH

Slide 28: Salt Hydrolysis

- Salts can be acidic or basic
- pH of a ____________ reaction depends on the extent that the ions (conjugate acids/bases) can hydrolyze water to form H^+ and OH^-
- Strength of conjugate acids and bases of the salt ____________ extent of hydrolysis and pH of solution
- **Types:**
- Salt of weak acid and strong base: anion ____________ (basic)
- Salt of strong acid and weak base: cation ____________ (acidic)
- Salt of weak acid and weak base: ____________ on K_a/K_b
- Salt of strong acid and strong base: ____________

- **Parent base + Parent acid → Salt + Water**
- **MOH HA M^+A^- H_2O**

Slide 29: Salt Hydrolysis

- **Salt of strong acid and weak base: cation ____________ (acidic)**
- **$M^+(aq) + H_2O(l) \rightleftharpoons MOH(aq) + H^+(aq)$**
- **$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_4OH(aq) + H^+(aq)$**
- Cation is a ____________ acid of the parent base (nonmetal)
- Release of H^+ causes pH of ____________ to decrease (pH < 7 at 298K)
- Metal cations with high charge ____________ (e.g. Al^{3+}, Fe^{3+}) can do the same
- **Salt of weak acid and strong base: anion ____________ (basic)**
- **$A^-(aq) + H_2O(l) \rightleftharpoons HA(aq) + OH^-(aq)$**
- **$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOOH(aq) + OH^-(aq)$**
- Anion is ____________ base of parent acid
- Release of OH^- causes pH of ____________ to increase (pH > 7 at 298K)

Slide 30: Salt Hydrolysis

- **Salt of weak acid and weak base: ____________ on K_a/K_b**
- Weak acids react with weak bases to form salts with strong ____________
- pH of ____________ depends on the K_a and K_b values of acids/bases

Slide 31: Summary

Slide 32: Let's Practice!

- Predict for each salt in aqueous solution whether the pH will be greater than, less than, or equal to 7.
- $NaCl$
- $FeCl_3$
- NH_4NO_3
- Na_2CO_3

Slide 33: Acid/Base Titrations

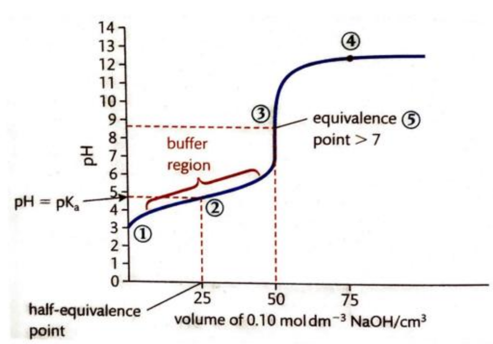
- Controlled volumes of one reactant added from a burette to a fixed volume of other reactant.
- Reaction takes place until equivalence point (stoichiometric point) is reached and neutralization occurs.
- Change in pH is not a linear relationship
- Values of pH is plotted against volume of reactant added to make a pH curve.
- A large jump in pH occurs at the **inflection point**. The equivalence point is half-way up the jump in inflection.
- pH of neutralized solution depends on hydrolysis by ions in salt.
- **In example problems, we will assume:**
- 0.10 mol dm^{-3} solutions all acids and bases
- Initial volume of 50.0 cm^3 of acid in Erlenmeyer flask (base in burette)
- Acids and bases react in 1:1 ratio (equivalence achieved when equal volumes of base added to acid).

Slide 35: Strong Acid/Strong Base Titrations

- 1. Initial pH = 1
- 2. pH changes _____________ until
- equivalence
- 3. Sharp jump in pH at _____________
- (pH 3 - 11)
- 4. Curve flattens after _____________
- 5. pH at _____________ = 7
- **$HCl(aq)$ + $NaOH(aq)$ → $NaCl(aq)$ + $H_2O(l)$**
- **Equivalence point: _____________ equal amounts of acid and base have neutralized each other**

Slide 36: Weak Acid/Strong Base Titrations

- 1. _____________ pH fairly high (weak acid)
- 2. pH stays relatively constant until _____________
- 3. Jump in pH at _____________ from pH 7.0 - 11.0 (lower span)
- 4. Curve flattens after _____________
- 5. pH > 7.0 at _____________
- **$CH_3COOH(aq)$ + $NaOH(aq)$ ⇌ $NaCH_3COO(aq)$ + $H_2O(l)$**
- **Half-_____________ point**

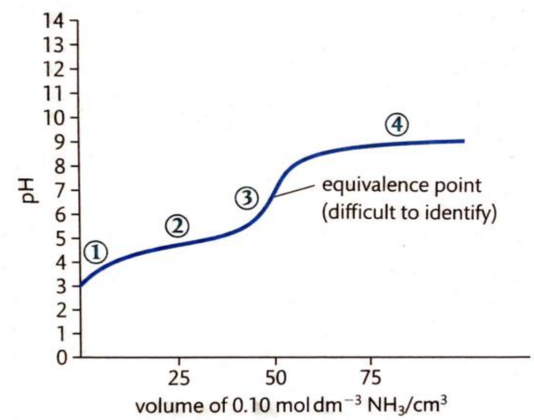


Slide 37: Strong Acid/Weak Base Titrations

- 1. ____________ pH = 1.0 (weak acid)
- 2. pH stays relatively constant until ____________
- 3. Jump in pH at ____________ from pH 3.0 - 7.0
- 4. Curve ____________ at low pH (pH of weak base)
- 5. pH $<$ 7.0 at ____________
- **$HCl(aq) + NH_3(aq) \rightleftharpoons NH_4Cl(aq)$**

Slide 38: Weak Acid/Weak Base Titrations

- 1. ____________ pH fairly high (weak acid)
- 2. ____________ of base causes pH to rise steadily
- 3. Change in pH at ____________ point is less sharp
- 4. Curve flattens after ____________ at fairly low pH (weak base)
- **$CH_3COOH(aq) + NH_3(aq) \rightleftharpoons NH_4CH_3COO(aq)$**
- **No clearly defined ____________ point**



Slide 39: Indicators

- Signal a change in pH by undergoing a color change.
- Weak acids/bases
- Undissociated and dissociated forms have different colors
- Change colors at the end-point when pH $= pK_a$
- Different indicators have different pKa values and have different end-points.
- Different indicator end-points are in **Section 22** of the Data Booklet
- Example of weak acid indicator:
- **HIn(aq)** $\rightleftharpoons H^+(aq) +$ **$In^-(aq)$**
- Increasing $[H^+]$: equilibrium shifts to the left
- Decreasing $[H^+]$: equilibrium shifts to the right
- $K_a =$
- $[H^+][In^-]$
- $[HIn]$
- At indicator equilibrium, $[In^-] = [HIn]$

- $K_a = [H^+]$ so $pK_a = pH$

Slide 40: It's important to use an indicator with a pKa near the equivalence point of your acid/base reaction

- **Indicators signal ____________ point**
- **End point is the pH at which a color change occurs**

Slide 41: Indicators

- How to choose an ____________:
- Determine what ____________ of weak and strong acid/base are reacting
- Deduce the pH of the salt solution at ____________
- Use an indicator with an end-point in the range of the ____________ point

Reactants in titration	pH range at equivalence	Example of suitable indicators	pK_a	End-point range of indicator and colour change
strong acid + strong base	3–11	phenolphthalein	9.50	8.2–10.0; colourless to pink
		methyl orange	3.46	3.2–4.4; red to yellow
weak acid + strong base	7–11	phenolphthalein	9.50	8.2–10.0; colourless to pink
		phenol red	8.00	6.6–8.0; yellow to red
strong acid + weak base	3–7	methyl orange	3.46	3.2–4.4; red to yellow
		bromophenol blue	4.10	3.0–4.6; yellow to blue
weak acid + weak base	this combination of acid and base does not give a significant change in pH at equivalence, so there is no suitable indicator to use here			

Slide 42: Let's Practice!

- Bromocresol green has a pH range of 3.8-5.4 and changes color from yellow to blue as the pH increases.
- Of the four types of titration shown in the table on slide 74, state in which two of these this indicator could be used
- Suggest a value for the pKa of this indicator
- What color will the indicator be at pH 3.6?