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DEPARTMENT OF THE INTERIOR
GEOLOGICAL SURVEY
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Dr. T. H. Johnson, Director
Division of Research
U. S. Atomic Energy Commission
16th St. and Constitution Ave., N. W.
Washington 25, D. C.

Dear Dr. Johnson:

Transmitted herewith is one copy of TEI-455, "Some thermodynamic relations among the uranium oxides, and their relation to the oxidation states of the uranium ores of the Colorado Plateaus," by Robert M. Garrels, August 1954.

We plan to submit this report for publication in the American Mineralogist.

Sincerely yours,

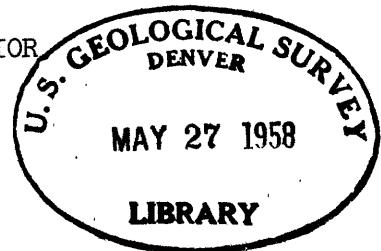
Andrew Brown
for W. H. Bradley
Chief Geologist

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Geology and Mineralogy

This document consists of 30 pages.
Series A.

UNITED STATES DEPARTMENT OF THE INTERIOR
GEOLOGICAL SURVEY



SOME THERMODYNAMIC RELATIONS AMONG THE URANIUM OXIDES
AND THEIR RELATION TO THE OXIDATION STATES OF THE
URANIUM ORES OF THE COLORADO PLATEAU*

By

Robert M. Garrels

August 1954

Trace Elements Investigations Report 455

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*This report concerns work done on behalf of the Division of Research of the U. S. Atomic Energy Commission.

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SOME THERMODYNAMIC RELATIONS AMONG THE URANIUM OXIDES
AND THEIR RELATION TO THE OXIDATION STATES OF THE
URANIUM ORES OF THE COLORADO PLATEAUS

By Robert M. Garrels

ABSTRACT

Fields of stability of uranium (VI) and uranium (IV) hydroxides and oxides in water solution at 25°C and one atmosphere pressure have been calculated as functions of Eh and pH. Equilibrium values of the activity of UO_2^{++} ion and of U^{4+} ion also have been calculated and are shown as contours on the stability fields. Thermodynamic relations among the uranium (VI) hydroxides and hydrated oxides indicate that the free energy differences among the various species are small. The data are interpreted to mean that a variety of such uranium (VI) compounds may form and even coexist. Similar studies of the uranium (IV) hydroxide indicate that it is unstable relative to the oxide and might well be expected to change to the oxide at a finite rate. Uranium (V) compounds probably have a transitory existence because of the instability of the UO_2^+ ion; uranium (III) oxides and hydroxides would not be expected to occur naturally because the uranium (III) ion would decompose water. A comparison of the behavior of the vanadium (III) and (IV) hydroxides with uranium (IV) oxide and uranium (VI) hydroxides indicates that vanadium (III) hydroxide should oxidize to the vanadium (IV) hydroxide at a lower potential than that required for the change from uranium (IV) to uranium (VI). A rather highly speculative diagram showing probable fields of stability of many of the major minerals of the Colorado Plateaus is

presented and the suggestion made that a consistent picture results if it is assumed that the ores, as viewed today, represent the superimposition of a weathering environment on a mineral assemblage formed in a primary reducing environment. It is emphasized that such an interpretation is consistent but not necessarily unique.

INTRODUCTION

Prior to about 1950, the uranium-vanadium ores of the Colorado Plateaus presented few problems in the interpretation of the oxidation-reduction conditions of their environment of formation. The oxidation-reduction "sensitive" elements of the major minerals uniformly exhibited the higher valences consistent with moderate to strongly oxidizing environments. During the last several years this consistency has disappeared; with the discovery of large quantities of "blue-black" or "black" ores, a variety of minerals containing the same elements, but in lower valence states, have become of economic importance.

The genetic relation of the "blue-black" ores to the more oxidized, or "carnotite" ores, is not clear, but one definite possibility is that the primary ores were an assemblage of minerals deposited under reducing conditions, and hence with the "sensitive" elements in their lower valence states, and that the more oxidized ores have been, at least in part, derived from them by ordinary weathering processes. If so, the primary oxidizing agent was oxygen carried by groundwaters, and the temperature and pressure of the reactions were close to 25 C. and 1 atmosphere.

It is possible to calculate the stability relations of many of the minerals involved under such conditions. In a previous paper (Garrels, 1953) relations of some of the vanadium oxides were considered. In this

report various uranium oxides, hydroxides, and hydrated oxides, for which thermodynamic data are available, are discussed.

This work was done on behalf of the Division of Research of the U. S. Atomic Energy Commission. C. R. Naeber of the Geological Survey was a constant source of information and guidance during the preparation of the manuscript and gave freely of his time in clarifying numerous aspects of the chemistry for the author.

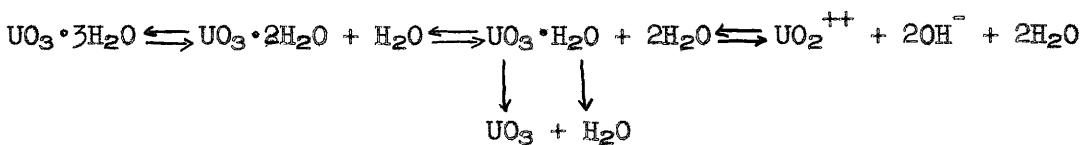
THERMODYNAMIC PROPERTIES OF SOME COMPOUNDS

AND IONS OF THE SYSTEM U-O-H₂O

Data on the thermodynamic properties of compounds and ions in the system U-O-H₂O have been assembled (table 1) from Latimer (1952) and from Seaborg and Katz (1954), with the exception of the bracketed values which have been calculated by the present author. The method of calculation is described in the appendix.

Relations among the uranium (VI) compounds

If UO₃·3H₂O is put into water, the following reactions among the known species are possible (coexistence at equilibrium of all the species listed is not implied):



The monohydrate is shown ionizing to the uranyl and hydroxide ions because of the likelihood that UO₃·H₂O is better expressed as UO₂(OH)₂. The free energy changes of the various hydration reactions UO₃·3H₂O \rightleftharpoons UO₃·2H₂O \rightleftharpoons UO₃·H₂O \rightleftharpoons UO₂⁺⁺ + 2OH⁻ + 2H₂O probably are small. This

Table 1.--Thermodynamic properties of some compounds and ions
of the system U-O-H₂O at 25°C.

(Heat and free energy of formation in kcal; entropy
of substance in cal/deg. Data from Latimer,
1952, except as noted).

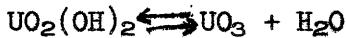
Compound or ion	ΔH°	ΔF°	S°	State
UO ₃ ·3H ₂ O		1/ [-376]		Crystalline
UO ₂ ·2H ₂ O [UO ₂ (OH) ₂ ·H ₂ O]	-446.2	1/ [-399]	1/ [40.4]	Crystalline
UO ₃ ·H ₂ O	-375.4	-343	(33)	Crystalline
UO ₃	2/-291.6	2/-273.1	23.57	Crystalline
UO ₂	2/-259.2	2/-246.6	18.6	Crystalline
U(OH) ₄		3/(-351.6)		Crystalline
U(OH) ₃		3/(-263.2)		Crystalline
UO ₂ ²⁺	-250.4	-236.4	-17	Aqueous
UO ₂ ⁺	-247.4	-237.6	12	Aqueous
U(OH) ³⁺	-204.1	-193.5	-30	Aqueous
U ⁴⁺	-146.7	-138.4	-78	Aqueous
U ³⁺	-123.0	-124.4	-30?	Aqueous
H ₂ O	- 68.317	- 56.690	16.716	Liquid
O ₂	0.0	0.0	49.003	Gaseous
H ₂	0.0	0.0	31.211	Gaseous
OH ⁻	- 54.957	- 37.595	- 2.52	Aqueous

1/ Estimated by the author

2/ Seaborg and Katz (1954, p. 174)

3/ Estimated by Latimer

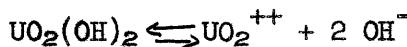
conclusion is based in part on analogy with the free energies of hydration of other compounds, and in part on the fact that ΔF° for the formation of $\text{UO}_2(\text{OH})_2 \cdot \text{H}_2\text{O}$, calculated from ΔH° and an estimated value of ΔS° , corresponds closely to the value calculated on the assumption that the free energy change of hydration of $\text{UO}_2(\text{OH})_2$ is zero. (See appendix.) On the other hand, UO_3 seems definitely unstable relative to $\text{UO}_2(\text{OH})_2$:



$$\Delta F^{\circ} = \Delta F^{\circ}_{\text{UO}_3} + \Delta F^{\circ}_{\text{H}_2\text{O}} - \Delta F^{\circ}_{\text{UO}_2(\text{OH})_2}$$

$$\Delta F^{\circ} = -273.1 + (-56.7) - (-343) = 13.2 \text{ kcal.}$$

There is also a large free energy change accompanying the ionization of $\text{UO}_2(\text{OH})_2$:



$$\Delta F^{\circ} = \Delta F^{\circ}_{\text{UO}_2^{++}} + 2 \Delta F^{\circ}_{\text{OH}^-} - \Delta F^{\circ}_{\text{UO}_2(\text{OH})_2}$$

$$\Delta F^{\circ} = -236.4 + (-75.2) - (-343) = +31.4 \text{ kcal.}$$

From this free energy change the activity product of $\text{UO}_2(\text{OH})_2$ can be calculated:



$$a_{\text{UO}_2^{++}} a_{\text{OH}^-}^2 = K_{\text{UO}_2(\text{OH})_2}$$

$$\Delta F^{\circ} = -RT \ln K$$

$$\text{At } 25^\circ \text{C, } \Delta F^{\circ} \text{ (kcal)} = -1.364 \log K \text{ (Latimer, 1952, p. 8).}$$

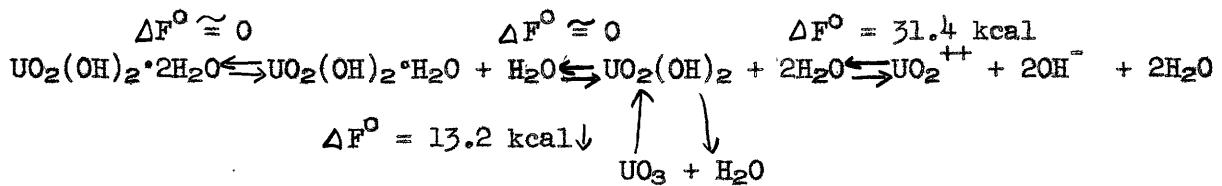
Therefore:

$$31.4 = -1.364 \log a_{\text{UO}_2^{++}} a_{\text{OH}^-}^2$$

$$a_{\text{UO}_2^{++}} a_{\text{OH}^-}^2 = 10^{-23}$$

..... (1)

From these relations, the free energy changes of the various reactions of uranium (VI) in water can be summarized diagrammatically:



Therefore it would seem that the various hydrated oxides have no very great tendency to transform one into the other, and that several might be expected to coexist metastably for considerable periods of time. Anhydrous UO_3 would not be expected, even as a metastable phase. $\text{UO}_2(\text{OH})_2$ does not ionize freely, and in the absence of complexes would dissolve perceptibly only in moderately acid solutions.

Relations among uranium (V) compounds

Uranium (V) can be formed metastably in water solution (Kraus, Nelson, and Johnson, 1949) as a more or less hydrolyzed UO_2^+ ion. Its rate of disproportionation into U^{4+} and UO_2^{++} ions, according to the reaction: $2\text{UO}_2^+ + 4\text{H}^+ \rightarrow \text{U}^{4+} + \text{UO}_2^{++} + 2\text{H}_2\text{O}$, is at a minimum between pH values of 2 and 4. The rate is finite, however, even in this range, so that the disproportionation may be a clue to the apparent lack of UO_2^+ compounds in nature.

Relations among the uranium (IV) compounds

In the system under consideration, free energy data are available only for the uranium (IV) compounds $\text{U}(\text{OH})_4$ and UO_2 . The reactions with water are:



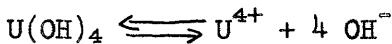
The free energy change of the reaction on the left is:

$$\Delta F^\circ = \Delta F^\circ_{U(OH)_4} - \Delta F^\circ_{UO_2} - 2\Delta F^\circ_{H_2O}$$

$$\Delta F^\circ = -351.6 - (-246.6) - (-113.4) = +8.4 \text{ kcal.}$$

Therefore, the hydroxide is definitely unstable with respect to the oxide, and, although the free energy relations give no information on rates, the hydroxide would not be expected to form and persist indefinitely, but to change to the oxide. The instability of the uranium (IV) hydroxide is in accord with the instability of Am (IV), Th (IV), and Sn (IV) hydroxides. The hydroxide, however, probably forms metastably during any laboratory precipitation, and then changes to the oxide on standing, so that the measured solubility would be that of the hydroxide, rather than that of the oxide. The probable relations between oxide and hydroxide can be shown from a consideration of their activity products.

For the ionization of $U(OH)_4$:



$$\Delta F^\circ = \Delta F^\circ_{U^{4+}} + 4\Delta F^\circ_{OH^-} - \Delta F^\circ_{U(OH)_4}$$

$$\Delta F^\circ = -138.4 + (-150.4) - (-351.6) = +62.8 \text{ kcal.}$$

From this value of ΔF° , the activity product of $U(OH)_4$ is:

$$a_{U^{4+}} a_{OH^-}^4 = K_{U(OH)_4}$$

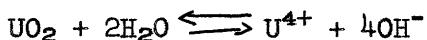
$$\Delta F^\circ = -RT \ln K$$

$$62.8 = 1.364 \log a_{U^{4+}} a_{OH^-}^4$$

$$a_{U^{4+}} a_{OH^-}^4 = 10^{14.6} = K_{U(OH)_4}$$

.... (2)

For the direct ionization of UO_2 :



$$\Delta F^\circ = \Delta F^\circ_{U^{4+}} + 4\Delta F^\circ_{OH^-} - \Delta F^\circ_{UO_2} - 2\Delta F^\circ_{H_2O}$$

$$\Delta F^\circ = -138.4 + (-150.4) - (-246.6) - 2(-56.7) = +71.2 \text{ kcal.}$$

The activity product of UO_2 is:

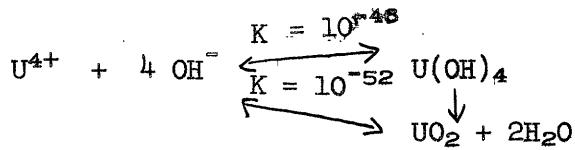
$$a_{\text{U}^{4+}} a_{\text{OH}^-} = K_{\text{UO}_2}$$

$$\Delta F^\circ = -RT \ln K_{\text{UO}_2}$$

$$71.2 = 1.364 \log K_{\text{UO}_2}$$

$$K_{\text{UO}_2} = a_{\text{U}^{4+}} a_{\text{OH}^-}^4 = 10^{-52}$$

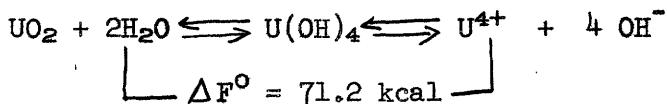
When U^{4+} is precipitated from water solution, two paths seem possible:



If precipitation takes place at true equilibrium, the precipitate should be UO_2 , which has the smaller K . But experimentally it is almost impossible to avoid exceeding both K_{UO_2} and $K_{\text{U(OH)}_4}$, and the metastable U(OH)_4 would be expected to form more rapidly, changing later to UO_2 .

In summary:

$$\Delta F^\circ = 8.4 \text{ kcal.} \quad \Delta F^\circ = 62.8 \text{ kcal}$$



Uranium (IV) would be expected to come out of solution as the hydroxide, and then change to the oxide, but the true equilibrium is between UO_2 and the ions.

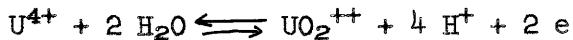
Relations among the uranium (III) compounds

It seems unlikely that uranium (III) compounds would be found under natural conditions at low temperature (Latimer, 1952, p. 301), for the U^{3+} ion is highly unstable with respect to the liberation of hydrogen from

water and would be expected to react immediately to a higher valence.

Relations between the uranium (VI) and the
uranium (IV) systems

In water solution, the uranium (VI) and uranium (IV) compounds can be considered to be connected through the UO_2^{++} and the U^{4+} ions. Under oxidizing conditions there is a tendency to transform U^{4+} to UO_2^{++} , and as a result to shift the equilibria toward the formation of uranium (VI) compounds. Under reducing conditions the reverse is true. The relation can be expressed as an equation:



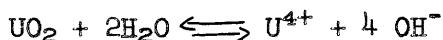
The free energy change can be calculated as before:

$$\Delta F^\circ = \Delta F^\circ_{\text{UO}_2^{++}} + 4 \Delta F^\circ_{\text{H}^+} - \Delta F^\circ_{\text{U}^{4+}} - 2 \Delta F^\circ_{\text{H}_2\text{O}}$$

$$\Delta F^\circ = -236.4 + (0) - (-138.4) - (-113.4) = +15.4 \text{ kcal.}$$

In other words, at unit activity of all constituents, UO_2^{++} tends to reduce to U^{4+} .

Because the uranium (VI) system presumably goes into solution through the ionization of $\text{UO}_2(\text{OH})_2$, and the uranium (IV) system through UO_2 , the three controlling reactions are:



The equilibrium relations are:

$$a_{\text{UO}_2^{++}} a_{\text{OH}^-}^2 = 10^{-23} \quad (1)$$

$$a_{\text{U}^{4+}} a_{\text{OH}^-}^4 = 10^{-52} \quad (2)$$

$$\text{Eh} = -E^\circ + 0.03 \log \frac{a_{\text{UO}_2^{++}} a_{\text{H}^+}^4}{a_{\text{U}^{4+}}} \quad (3)$$

Equation 3 is the standard oxidation equation relating oxidation potential of the system (E_h) to the standard electrode potential (E°) and the ionic ratios. Pourbaix (1949) gives a fuller discussion of this relation.

If $UO_2(OH)_2$ and UO_2 are in equilibrium, equations 1 and 2 hold. Then, dividing 1 by 2, and transforming:

$$\frac{a_{UO_2^{++}}}{a_{U^{4+}}} = 10^{29} a_{OH^-}^2 \quad (4)$$

and:

$$\log \frac{a_{UO_2^{++}}}{a_{U^{4+}}} = 29 + 2 \log a_{OH^-} \quad (5)$$

The a_{OH^-} can be related to the pH. Because

$$a_H^+ a_{OH^-} = K_w = 10^{-14} \text{ at } 25^\circ\text{C.}$$

then:

$$a_{OH^-} = \frac{K_w}{a_H^+}$$

and:

$$\log a_{OH^-} = \log K_w - \log a_H^+$$

By definition, $pH = -\log a_H^+$, so:

$$\log a_{OH^-} = pH - 14$$

Substituting this relation in equation 5:

$$\log \frac{a_{UO_2^{++}}}{a_{U^{4+}}} = 29 + 2(pH - 14)$$

or:

$$\log \frac{a_{UO_2^{++}}}{a_{U^{4+}}} = 2pH + 1 \quad (6)$$

Equation 3 can be rewritten to contain pH:

$$E_h = -E^{\circ} + 0.03 \log \frac{a_{UO_2^{++}}}{a_{U^{4+}}} - 0.12 \text{ pH}$$

Substituting the value from equation 6 for $\log \frac{a_{UO_2^{++}}}{a_{U^4+}}$:

$$Eh = -E^\circ + 0.03 (2 \text{ pH} + 1) - 0.12 \text{ pH}$$

and:

$$Eh = -E^\circ + 0.03 - 0.06 \text{ pH}$$

Therefore, Eh is a linear function of pH at conditions of equilibrium for UO_2 and $UO_2(OH)_2$.

Substituting the numerical value of E° , which is -0.334 volts at 25° (Latimer, 1952, p. 301), the simple equation is obtained:

$$Eh = 0.364 - 0.06 \text{ pH} \quad (7)$$

Figure 1 shows the boundary between the stability fields of the two compounds, calculated from equation 7. Above the boundary $UO_2(OH)_2$ is the stable solid phase; below, UO_2 . Because of the small free energy differences between the various uranium (VI) hydrates, it is not unlikely that any one or a combination of them might form and persist. On the other hand, the uranium (IV) oxide is definitely stable with respect to the hydroxide, which, if formed, would be expected to change at a finite rate to UO_2 . The boundary shown for the stability of water apparently precludes the possibility of occurrence of uranium (III) compounds; the stability field of $U(OH)_3$ is more than 0.5 volts below this boundary, which practically assures that $U(OH)_3$ would decompose water at a finite rate.

Figure 2 shows the same stability fields with contours of the activity of UO_2^{++} ion superimposed.^{1/} The activity is significant ($> 10^{-6}$) only in mildly oxidizing acid solutions.

^{1/} For details of the construction of such contoured diagrams, see Krumbein and Garrels (1952) or Pourbaix (1949).

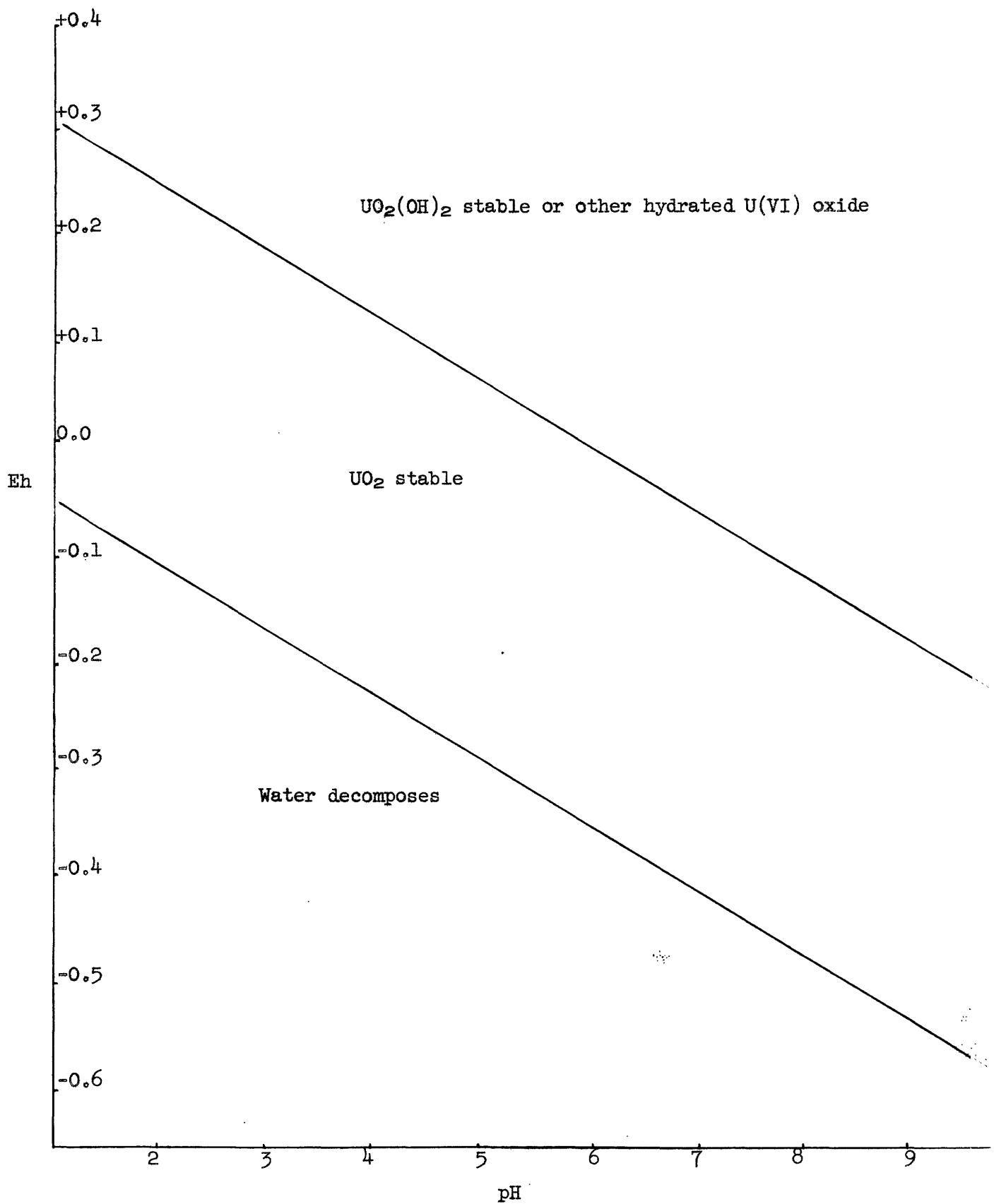


Figure 1.--Fields of stability of UO_2 and of hydrated U^{+6} oxides.

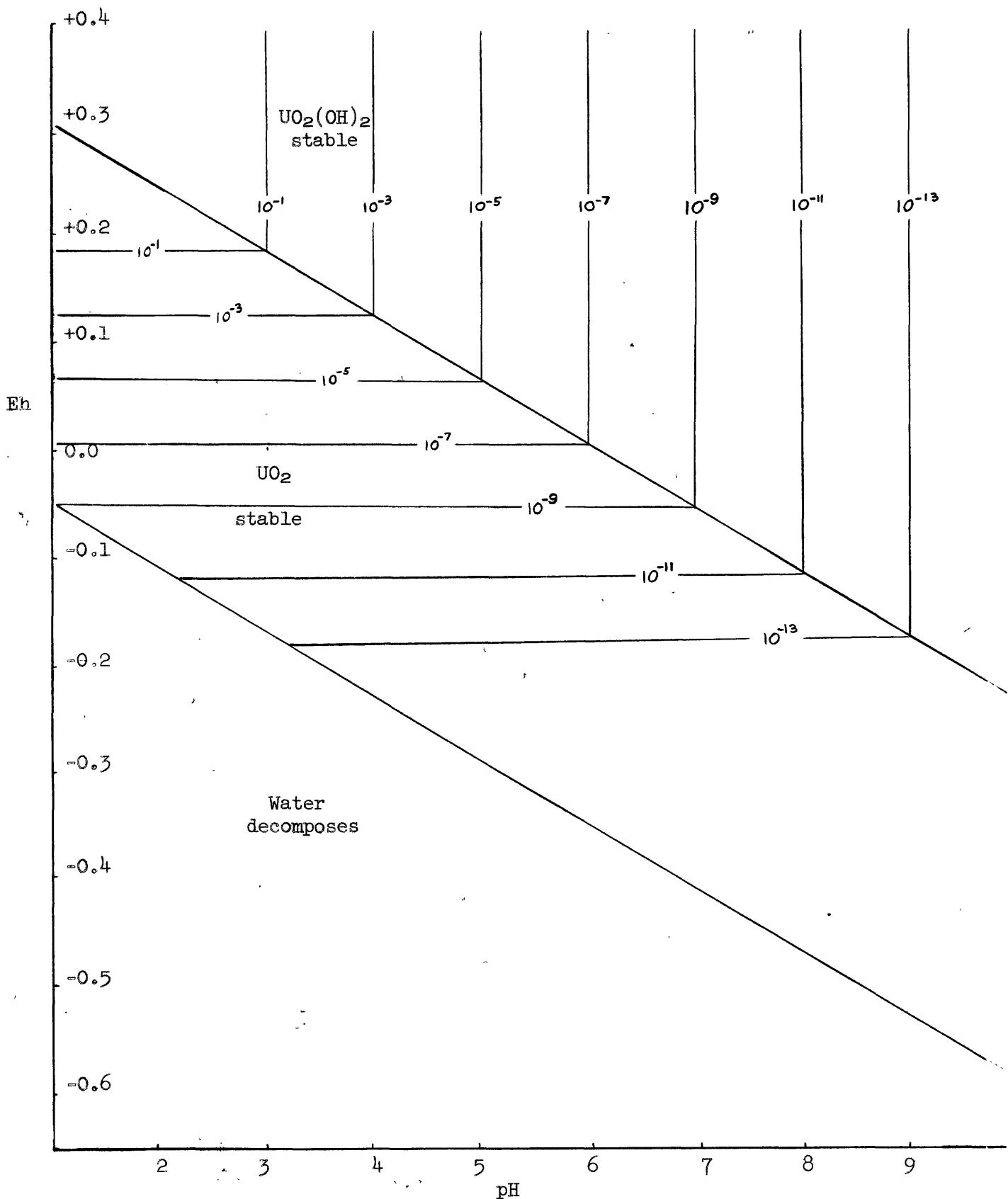


Figure 2.-- UO_2^{++} ion activity in equilibrium with $\text{UO}_2(\text{OH})_2$ and UO_2 .

In figure 3 contours of the activity of U^{4+} ion are shown. Activity exceeds 10^{-6} only at pH values of 2.5 or less.

Figure 4 shows contours of the sum of $a_{U^{4+}}$ and $a_{UO_2^{++}}$, and re-emphasizes the low activity of either ion in very weakly acid or alkaline solutions.

Figure 5, a plot of the ratio of UO_2^{++} activity to U^{4+} activity, shows that the contribution of U^{4+} ion is relatively insignificant except under acid reducing conditions. In other words, UO_2^{++} ion activity far exceeds that of U^{4+} under most conditions, even in the field of stability of UO_2 .

Relation of vanadium and uranium oxide stability fields

A study of the fields of stability of the vanadium oxides already has been made (Garrels, 1953). In figure 6 the fields of the uranium and vanadium oxides are shown together. Contours have been drawn to show conditions under which the total activity of various vanadium ions $\left[\sum a_{V^{++}} + a_{V^3+} + a_{VO^{++}} + a_{V(OH)_4^+} \right]$ is appreciable (10^{-2}) and negligible (10^{-6}). Similar treatment has been accorded $\sum a_{UO_2^{++}} + a_{U^{4+}}$.

DISCUSSION AND INTERPRETATION

It is possible to draw a theoretical and partly hypothetical diagram showing over-all Eh-pH stability relationships for an aqueous system containing the chief minerals of the various deposits on the Colorado Plateaus. (See fig. 7.)

Fields 1 to 5 on this diagram have reasonable validity in that they actually have been calculated from thermodynamic data for the chemical compounds. Fields 6 and 7 are very highly speculative. To delimit the

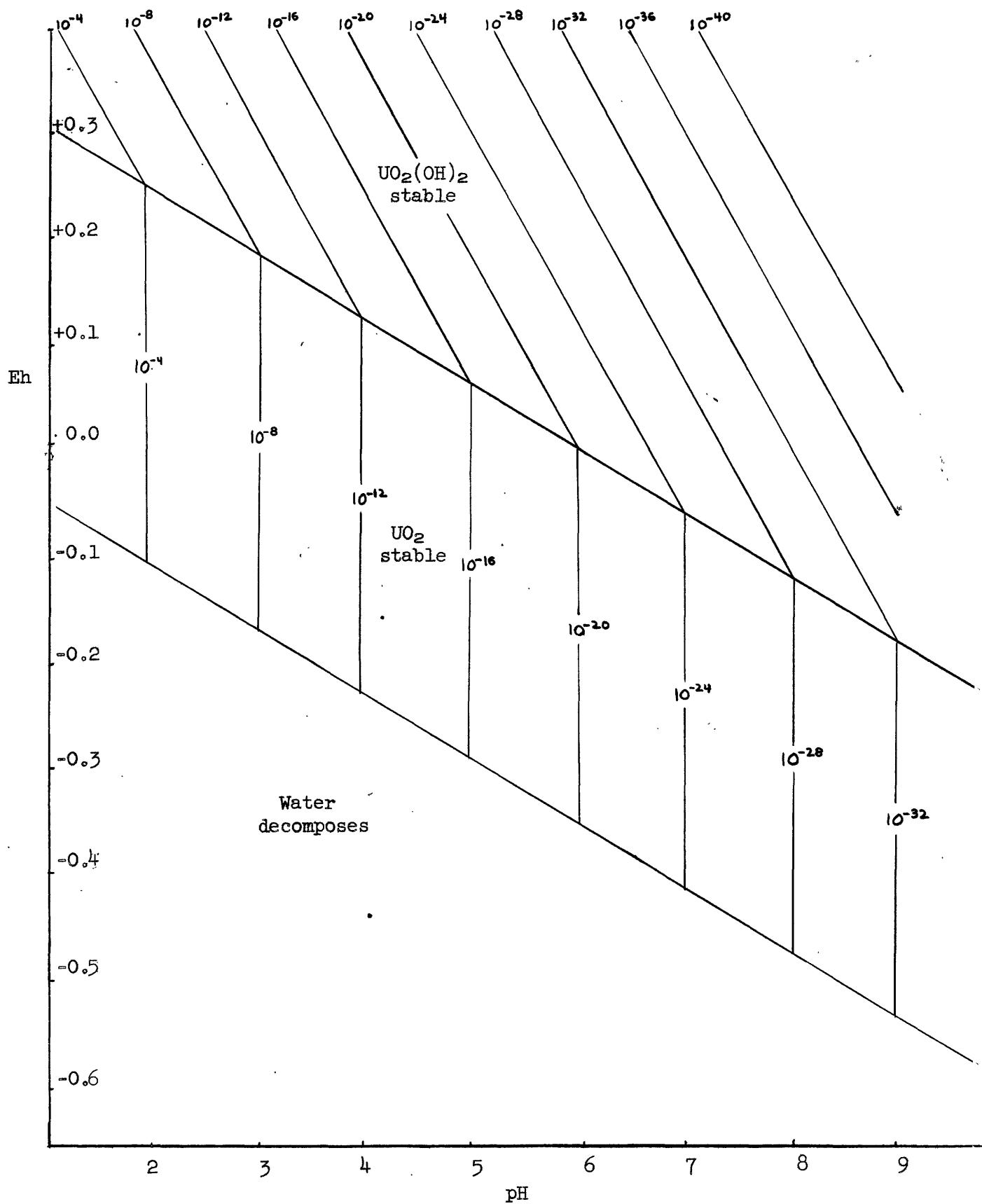


Figure 3.--U⁴⁺ ion activity in equilibrium with UO₂(OH)₂ and UO₂.

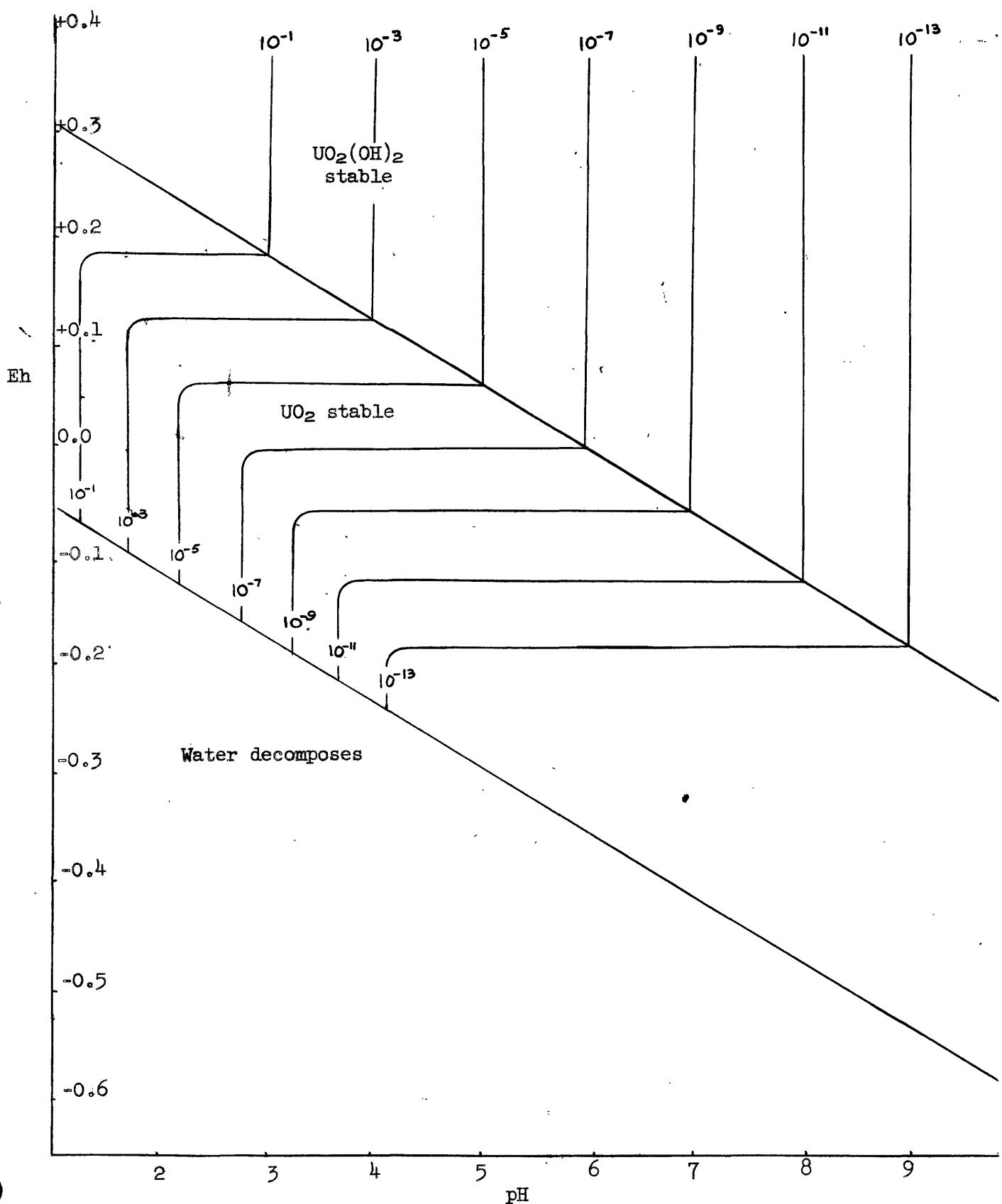


Figure 4. --Sum of activities of UO_2^{++} and U^{4+} ions in equilibrium with $\text{UO}_2(\text{OH})_2$ and UO_2

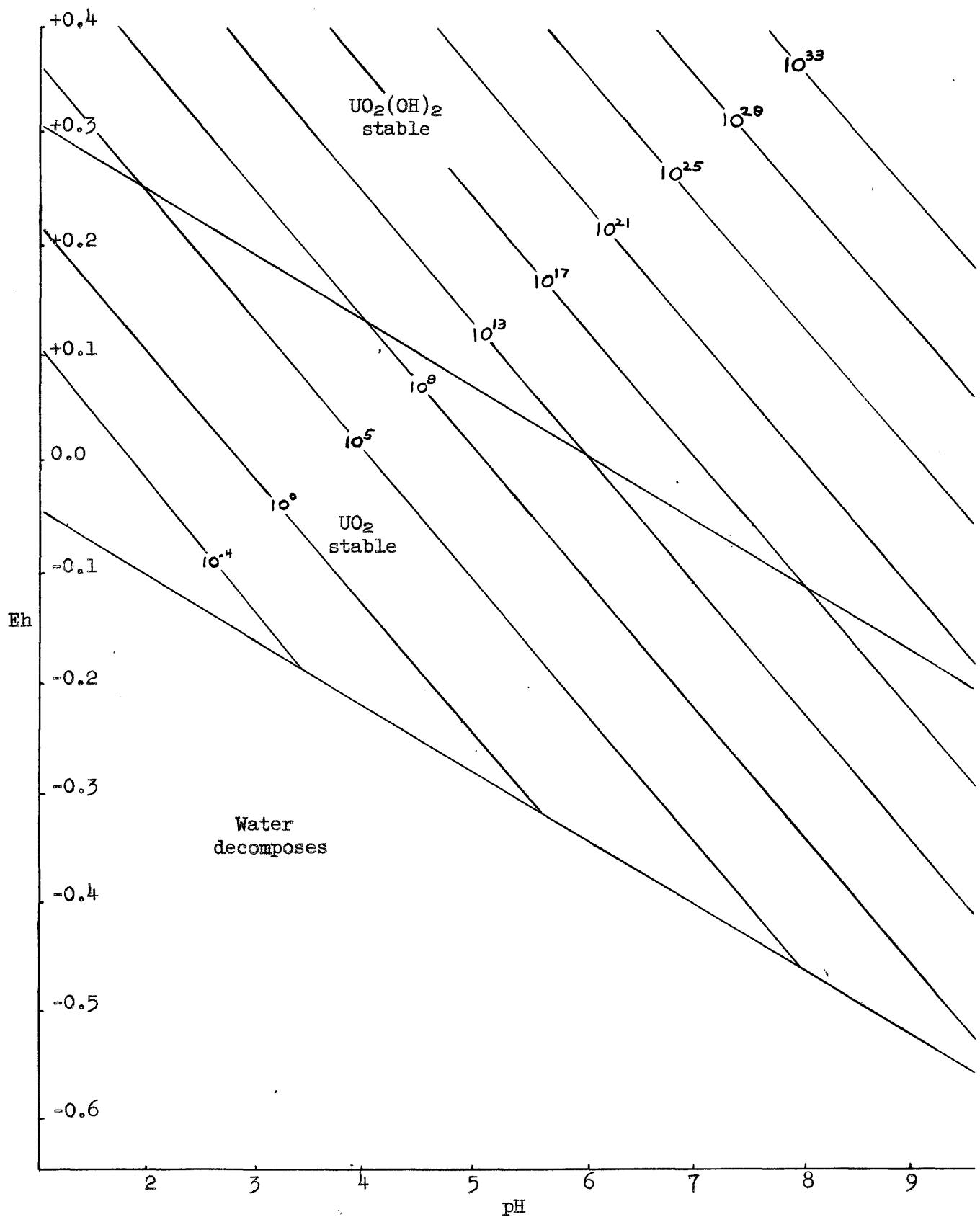


Figure 5.--Ratio of UO_2^{++} and U^{4+} ions in equilibrium with $\text{UO}_2(\text{OH})_2$ and UO_2 . Contours are of $\log \frac{a_{\text{UO}_2^{++}}}{a_{\text{U}^{4+}}}$

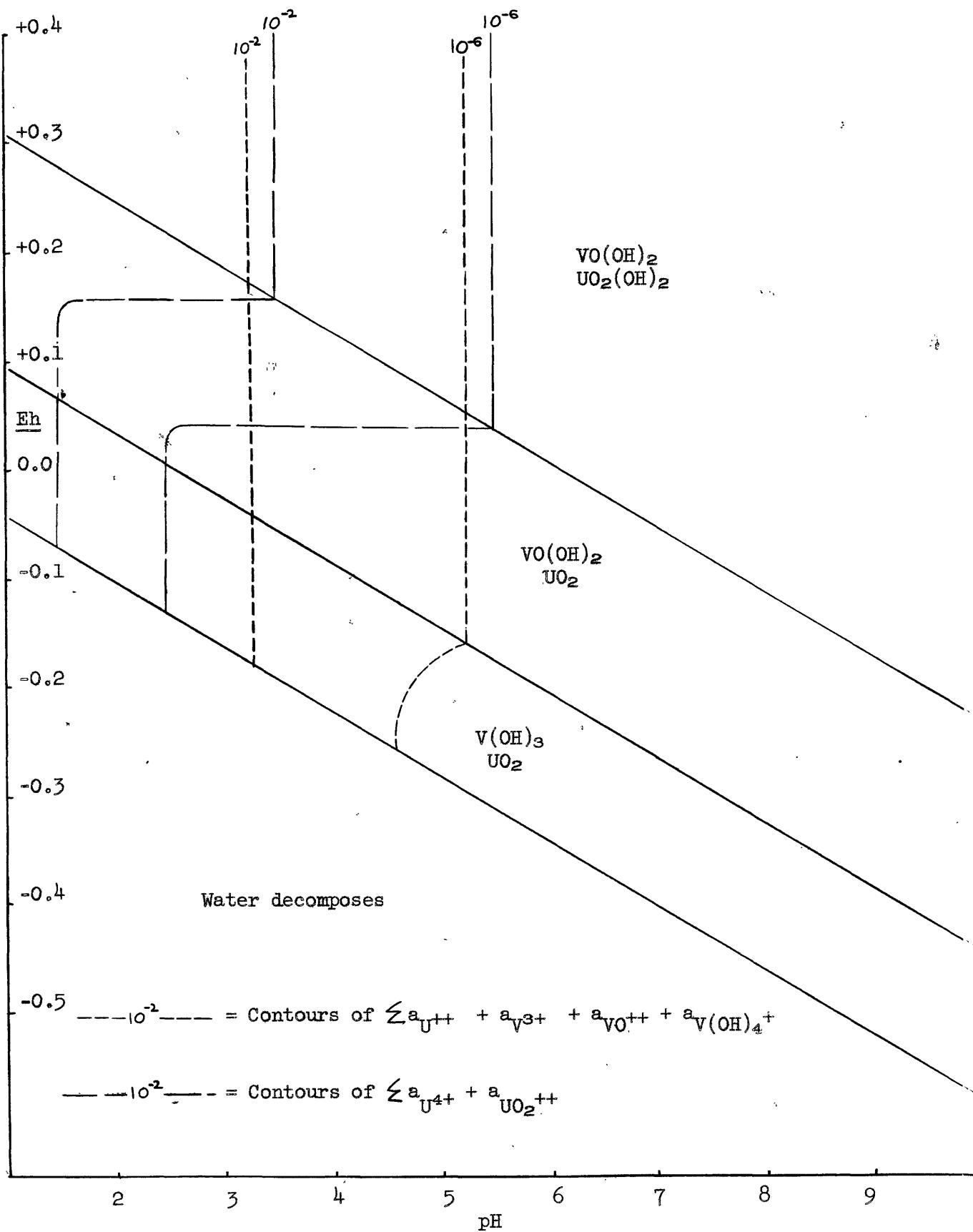


Figure 6.--Composite diagram showing fields of stability of some uranium and vanadium oxides and hydroxides.

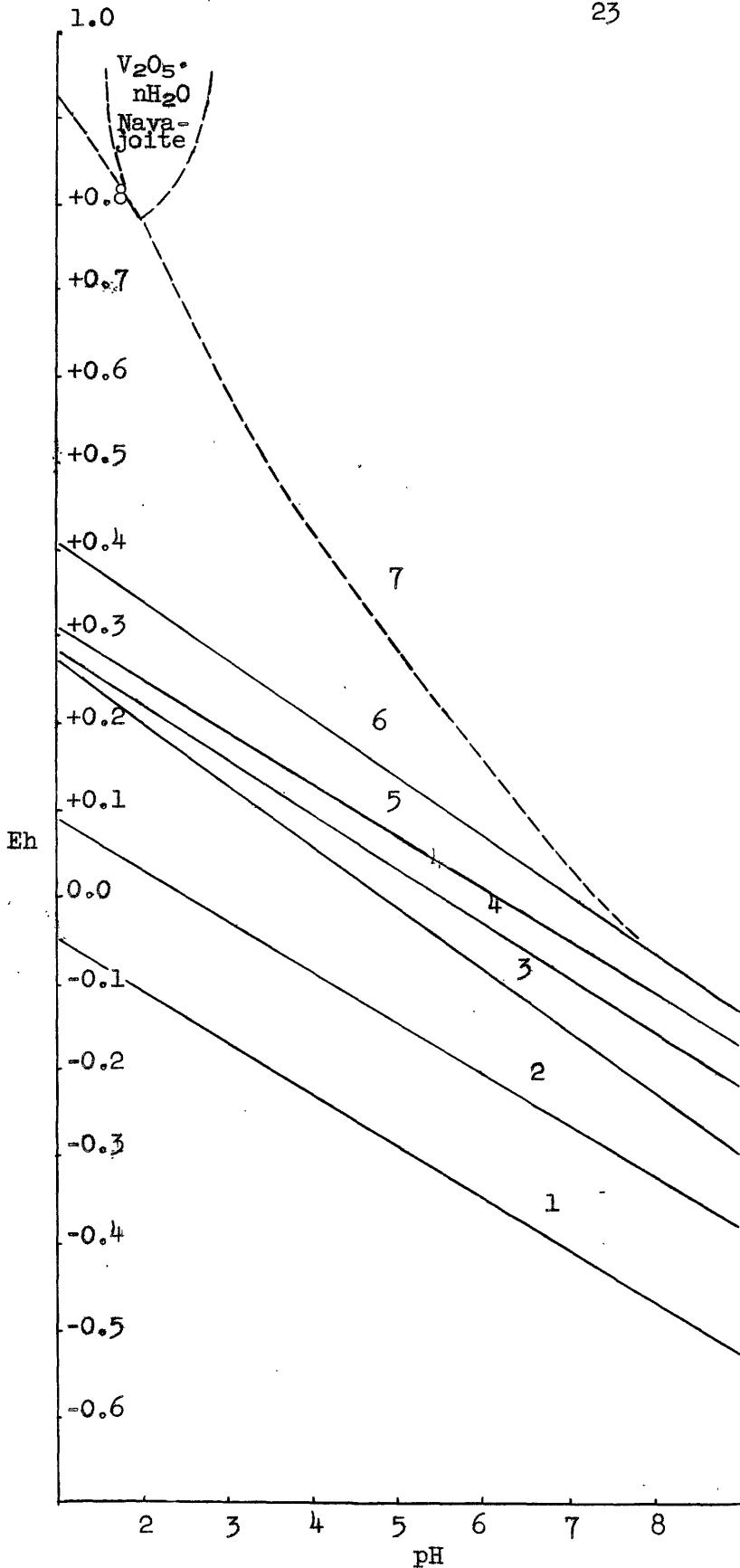


Figure 7.--Summary of information on stability relations of some major minerals of the Colorado Plateaus. Temp. = 25°C.; total sulfur ($S^{2-} + HS^- + H_2S$) = 0.1 M; total CO_2 $\Sigma CO_3^{2-} + HCO_3^- + H_2CO_3$ = 0.01 M. The boundaries of fields 6 and 7 are especially speculative.

stability fields of the various sulfides it has been necessary to make an assumption as to composition of the solutions. The assumption is that total sulfur ($\text{SO}_4^{=}$ + H_2S + HS^- + S^{\equiv}) is 0.1 mol per liter; this seems not unreasonable in that it is of the order of magnitude of total sulfur in mine water, which should be somewhat comparable to those of a naturally oxidizing ore deposit. Furthermore, a decrease of total sulfur to 0.001 molar would not change the pattern significantly.

The various fields delimited by the boundaries represent the best interpretation possible at the present time as to stable mineral assemblages. Perhaps the best device to illustrate the kinds of changes expected as a result of oxidation of the most reduced assemblage, is to discuss the progressive changes expected as the oxidation potential is increased, carrying this original assemblage upward through the various fields.

Field 1 is characterized by vanadium (III) oxides or hydroxides; uranium (IV) oxide; zinc, lead, copper, and iron sulfides. Mineralogically this might represent a mineral association of montroseite, uraninite, galena, sphalerite, covellite, and pyrite. The first of these to oxidize would be montroseite, and a field (field 2) presumably exists in which a vanadium (IV) oxide is stable in association with uraninite and the metal sulfides. Further increase in the oxidation of the environment through fields 3, 4, and 5 up into field 6, presumably would result in the oxidation in sequence of sphalerite, galena, pyrite, uraninite, and copper sulfide. In essence, then, the upper boundary of field 5 would seem to represent the most oxidizing conditions under which sulfide ion could resist oxidation up to sulfate. Zinc sulfide dissociates in water solution more easily than lead sulfide, which in turn dissociates more easily than copper sulfide; this explains the sequential oxidation of the sulfides, with sphalerite tending to decompose at lower Eh values. The exact position

of pyrite is not known, although studies of the relative rate of oxidation (Koch and Grasselly, 1951) suggest that its ease of oxidation, at least, is comparable to that of other sulfides.

Above field 5 the relations are highly hypothetical and no good data are known for the actual calculation of the stabilities of any of the uranium-vanadium compounds. Relations among some of the oxidized copper and lead compounds are known and the fields of stability have been delimited (Garrels, 1954), but it seems unnecessary to attempt to show detailed relationships among them because their uranium-vanadium counterparts have not been worked out. The dashed line separating fields 6 and 7, however, represents a rough boundary interpreted from the work of Ducret (1951, p. 729-737) on the vanadium (IV)-(V) oxidation potential. Field 7 is the area in which vanadium exists as a complex series of vanadate ions, and where the ratio of these ions, carrying vanadium (V), to ions with quadrivalent vanadium (VO^{++}), is large (greater than 10). Field 6 on the other hand is an area in which the ratio of vanadium (IV) to vanadium (V) is significant and an area in which Ducret (1951, p. 729-737) found a whole series of vanadium IV-V oxides of variable composition. Therefore, it seems reasonable to assign to field 6 the complex mixture of natural vanadium (IV) and (V) oxides called corvusite, and to field 7 the various metal-uranyl vanadates and metal vanadates such as carnotite, tyuyamunite, hewettite, pascoite, rossite, and others. One special point should be made in passing, at a pH of approximately 2 vanadium changes from an anionic role to a cationic one; from vanadates to the vanadyl (VO_2^+) ion (Ducret, 1951, p. 714). In this general pH range, during the transition, hydrated V_2O_5 precipitates; therefore, it seems likely that navajoite, the naturally occurring V_2O_5 hydrate (Weeks, Thompson, and

Sherwood, 1954), is an excellent indicator of rather strongly acid oxidizing solutions.

On the basis, then, of the estimation of stability fields as shown in field 7, a "narrative" interpretation of the oxidation of a primary montrosoelite, uraninite, pyrite, sphalerite, galena, covellite ore can be attempted. Under ordinary ground-water conditions the environment is reducing and alkaline (Garrels, 1953, p. 1263-1264). On the other hand, atmospheric oxygen brought into the system is a strong oxidizing agent, and even traces of oxygen tend to raise the oxidation potential to plus values of several tenths of a volt. The effect is to superimpose a strongly oxidizing environment, represented perhaps by fields 6 and 7, on the original reduced-mineral assemblage. The tendency is to pull all the compounds up to highly oxidized species. If equilibrium were attained instantaneously, the original association would jump to the associations of fields 6 and 7. The oxygen, however, is probably supplied slowly during weathering and the rate of reaction with the various species would be expected to differ markedly. If considered entirely from the oxidation potential standpoint, that is, if the rates are entirely functions of the difference in potential between the environment and the upper limit of mineral stability, then it would be expected that the vanadium oxides would weather to vanadium (IV) compounds before the sulfides were significantly attacked and that uraninite should oxidize to uranium (VI) compounds when the sulfides change to sulfate, oxides, and carbonates. Actual experiment is necessary, however, to assess the relative effect of the potential versus the rate characteristics of the individual species.

When the sulfides become unstable, sulfide ion is converted to sulfate, and the freed metal ions become involved in a variety of reaction depend-

ing upon the relative stabilities of the various solids that can form with the chief anionic constituents of the oxidized solutions. Lead tends to form slightly soluble sulfates and carbonates; zinc to form moderately soluble carbonates and silicates; copper to form a variety of oxides, carbonates and hydroxy-carbonates, as well as native copper and copper silicate. The iron from pyrite tends to hydrolyze into hematite or various hydrated ferric oxides.

The vanadium, after it reaches the quadrivalent stage, probably oxidizes up through a complex series of vanadium (IV) and (V) oxides (corvusite). As more and more of the vanadium reaches the quinquevalent state it tends to go into vanadates. Any uranium (VI), possibly present briefly in the pre-vanadate stage as uranium (VI) oxides, tends to be converted into potassium or calcium uranyl vanadates. Vanadium in excess of the amount necessary to form these uranyl compounds would tend to precipitate as simple metal vanadates.

SUMMARY AND CONCLUSION

The present information on the fields of stability of the minerals of the Colorado Plateau-type deposits is quite consistent with an interpretation that the original ores were species containing "sensitive" elements in their lower valence states, and that the great complexity of the mineralogy, as now observed, can be attributed to the superimposition of oxidation, of various degrees of completeness, through the primary agency of atmospheric oxygen. It should also be clearly emphasized that the present picture may also be interpreted on the basis of the influence of a variety of original environments acting at the time

of deposition, or by various combinations of these two major conceptual schemes. A determination of the actual processes can be made by extensive study of the paragenesis of the minerals and of the geologic occurrence of the ores. According to the "weathering hypothesis", oxidation should have taken place only in those places accessible to atmospheric oxygen. It should not be difficult to assess its validity if field studies are made with the question in mind.

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APPENDIX

Calculation of free energy of formation
for uranium compounds

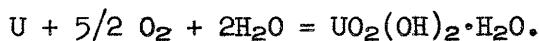
Two methods have been used to calculate the values in brackets in table 1. The first is by estimation of entropy values if ΔH° is known, and then using the relation $\Delta F^\circ = \Delta H^\circ - T \Delta S^\circ$; the second by using the assumption that the hydration reaction is essentially an equilibrium process, so that the free energies of formation of the products is equal to the sum of the free energies of formation of the reactants.

For $UO_3 \cdot 2H_2O$ both methods can be used. ΔH° is known; S° can be estimated by the methods outlined by Latimer (1952, p. 359-369).

Because uranium (VI) in water solution gives the stable UO_2^{++} ion, it seems likely that $UO_3 \cdot 2H_2O$ is better expressed as $UO_2(OH)_2 \cdot H_2O$. According to Latimer's values for the entropies of elements in compounds, uranium is 16, oxygen is 3, OH^- is 4.5, H_2O is 9.4.

U	16
2 O	6
20 OH^-	9
H_2O	9.4
$\overline{40.4} = S^\circ$	
$UO_2(OH)_2 \cdot H_2O$	

ΔS° can then be obtained from the reaction:



Using the entropies of the elements (table 1)

$$\Delta S^\circ = S^\circ_{UO_2(OH)_2 \cdot H_2O} - S^\circ_U - 5/2 S^\circ_{O_2} - 2S^\circ_{H_2}$$

$$\Delta S^\circ = 40.4 - 16.0 - 122.5 - 62.4 = -160.5$$

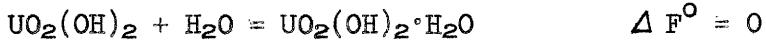
Then:

$$\Delta F^\circ = \Delta H^\circ - T \Delta S^\circ$$

At 25°C, ΔH° for the formation of $\text{UO}_2(\text{OH})_2 \cdot \text{H}_2\text{O}$ is -446.2 kcal (table 1). Then:

$$\Delta F^\circ = -446,200 - (298 \times -160.5) = 398,400 \text{ cal} = 398.4 \text{ kcal}$$

For the second method, the assumption is made:



$$\Delta F^\circ_{\text{UO}_2(\text{OH})_2} + \Delta F^\circ_{\text{H}_2\text{O}} = \Delta F^\circ_{\text{UO}_2(\text{OH})_2 \cdot \text{H}_2\text{O}}$$

From table 1:

$$-343 + (-56.7) = -399.7 \text{ kcal} = \Delta F^\circ_{\text{UO}_2(\text{OH})_2 \cdot \text{H}_2\text{O}}$$

Therefore the values by the two methods check within 1.3 kcal, or less than 0.4 percent difference.

For $\text{UO}_2 \cdot 3\text{H}_2\text{O}$ no value of ΔH° is available, and the bracketed value in table 1 is based on the second method:



$$\Delta F^\circ_{\text{UO}_2(\text{OH})_2} + \Delta F^\circ_{2\text{H}_2\text{O}} = \Delta F^\circ_{\text{UO}_2(\text{OH})_2 \cdot 2\text{H}_2\text{O}}$$

Substituting ΔF° values from table 1:

$$-343 + (-33.4) = -376.4 \text{ kcal} = \Delta F^\circ_{\text{UO}_2(\text{OH})_2 \cdot 2\text{H}_2\text{O}}$$